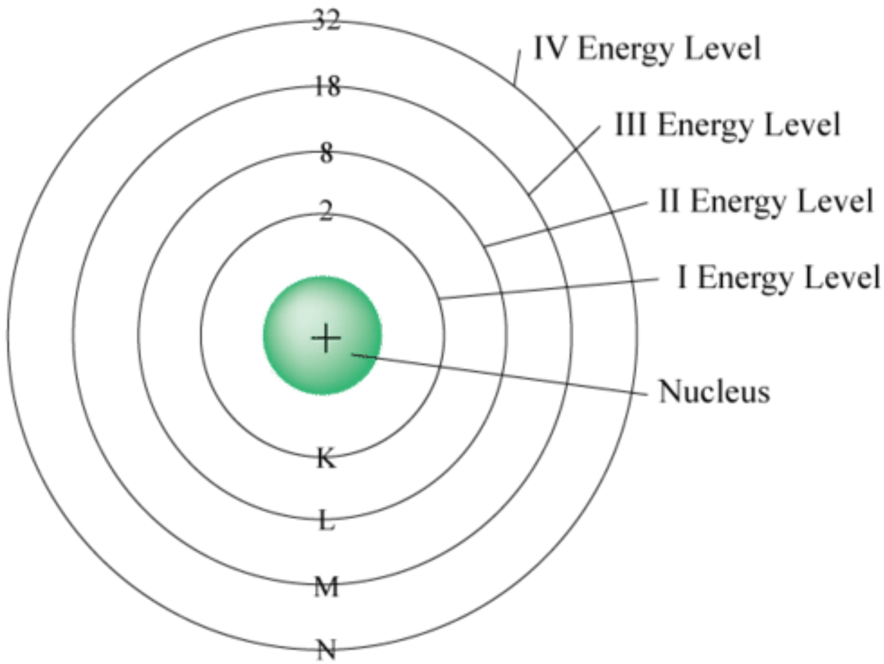
The Effects of Atomic Charges

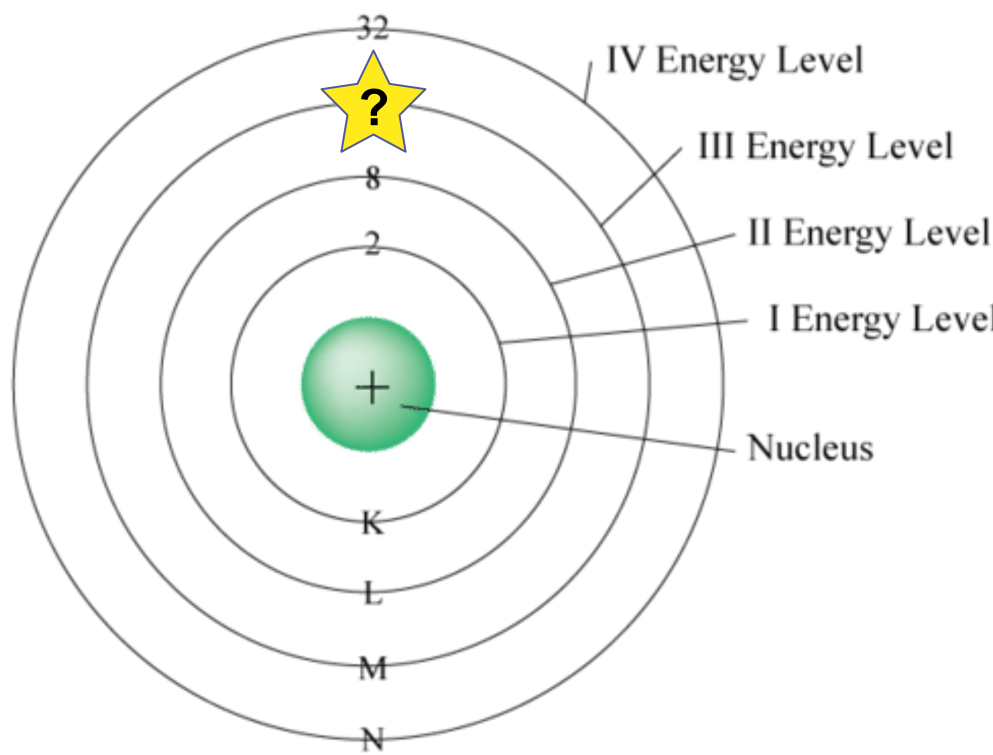
- **Protons in the nucleus attract electrons**
  - Electrons are located in the region of the atom knows as the electron cloud
    - Energy levels (also called electron shells) are fixed distances from the nucleus of an atom where electrons are most likely to be found.



  - The 1st shell or principal quantum number n = 1 is closest to the nucleus.
    - The 2nd energy level n = 2 follows n = 1. Then follows n = 3 etc.
  - Each level can theoretically hold a certain maximum number of electrons.
    - Maximum # of electrons in energy level n: X
    - Maximum # of electrons an energy level can hold: $X = 2n^2$

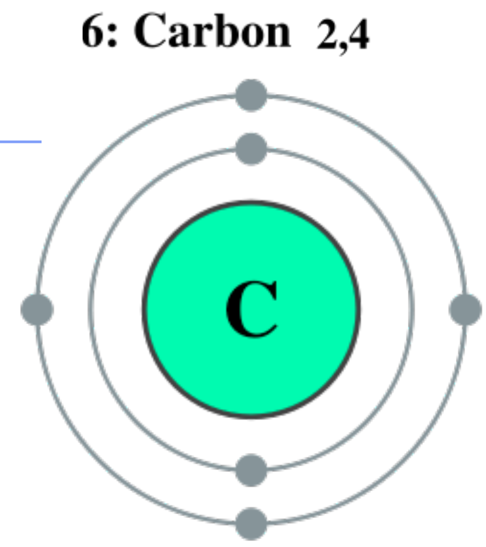
The Effects of Atomic Charges

- Maximum # of electrons an energy level can hold: $X = 2n^2$
  - EXERCISE: How many electrons can energy level 3 hold?
    - $X=2x^{32}=18$
  - Can this level hold 15 electrons?
    - Yes
  - What about 20?
    - No

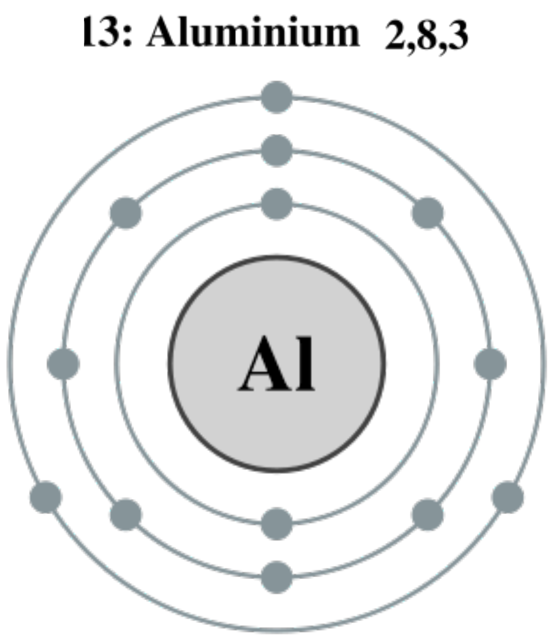


THE ELECTRON CLOUD

- Let's look at CARBON:
  - Atomic Number = 6 (6 p thus 6 e-).
    - The first 2 electrons fill shell n = 1.
    - The remaining 4 electrons fill shell n = 2.
      - Energy level 2 fills after 1 is full!

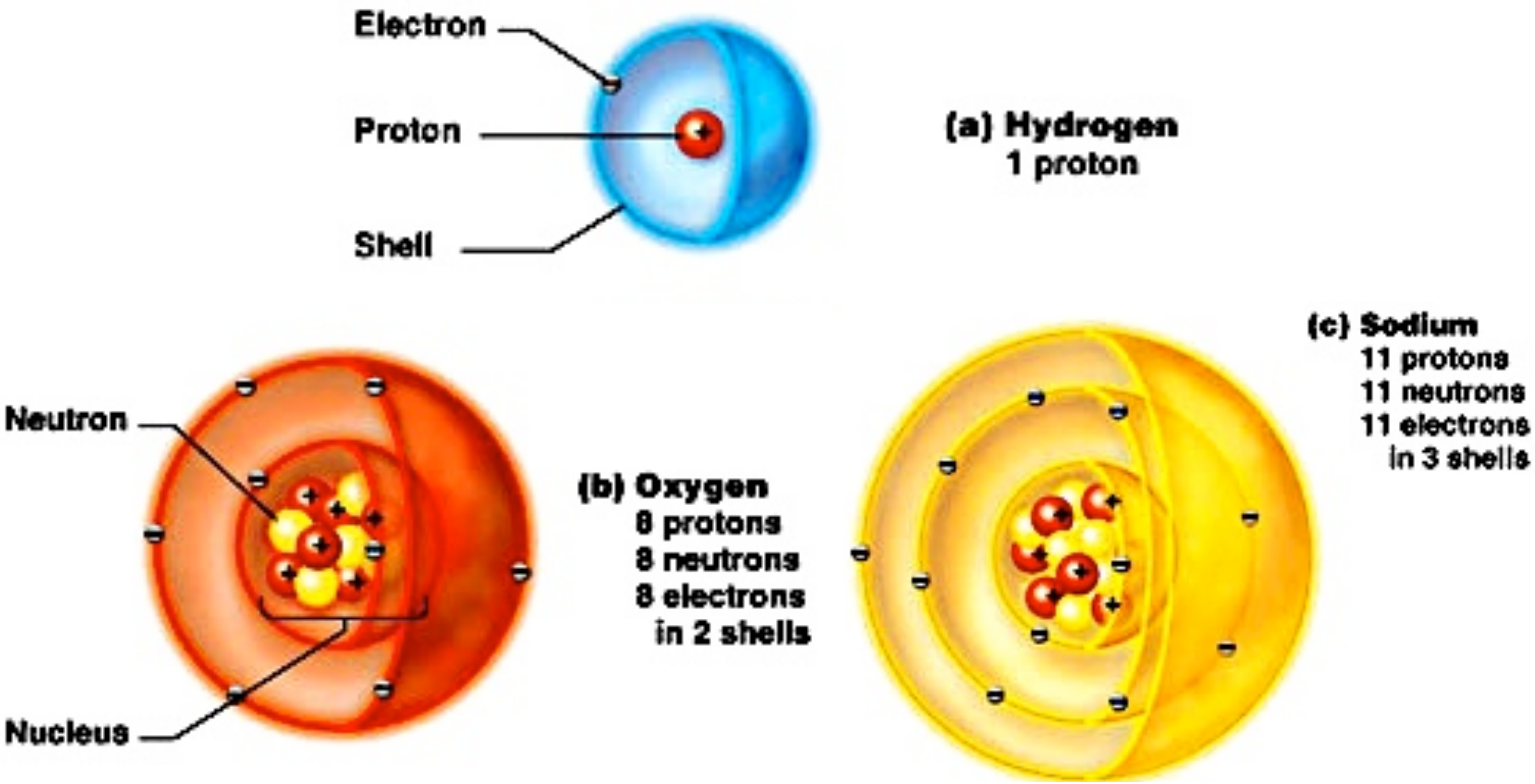


- Let's look at ALUMINUM:
  - Atomic Number = 13 (13 p thus 13 e-).
    - The first 2 electrons fill n = 1.
    - The next 8 electrons fill n = 2.
    - The last 3 electrons fill n = 3.
      - Energy level 3 fills after energy level 2 is full!



AP B

ATOMIC ENERGY LEVELS

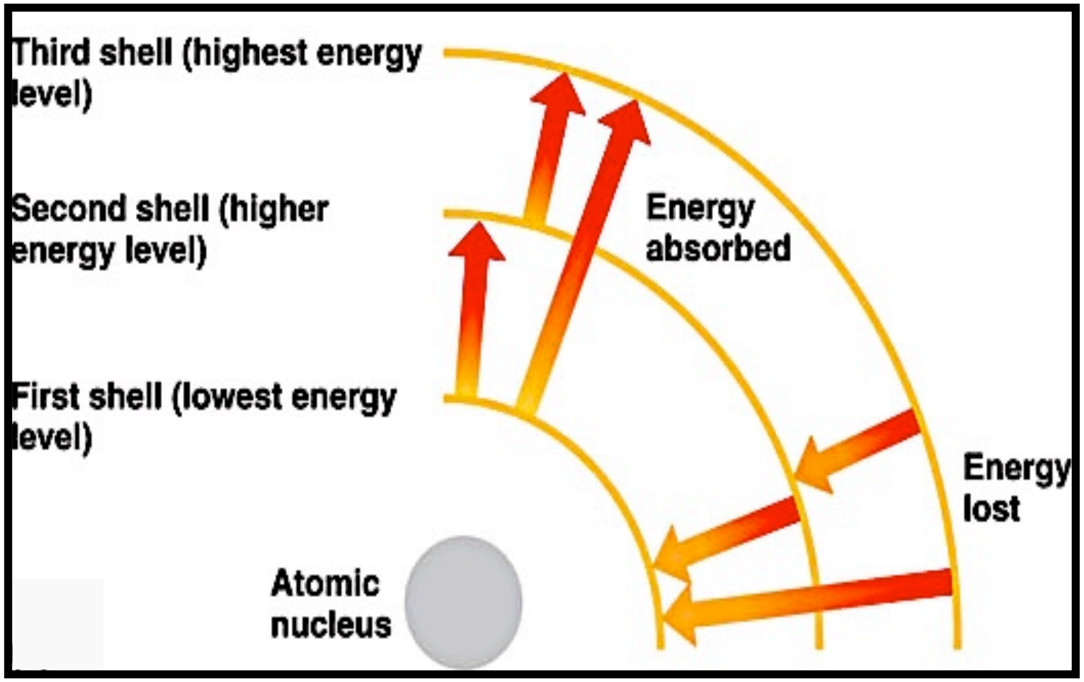


- Remember that the principal energy levels are more like "clouds" than shells. We can <u>never</u> be sure exactly where an electron is located but we can know where the electron will <u>most probably</u> be found.

AP

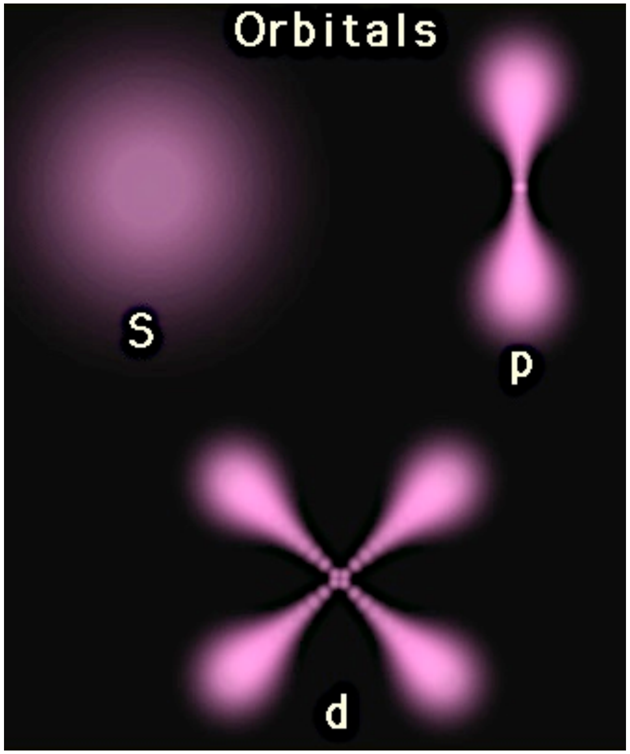
ATOMIC ENERGY LEVELS

- An electron sitting as close as possible to the nucleus has the lowest energy.
  - This atom is said to be in its lowest energy state or GROUND STATE.
- If a discrete quantity of additional energy were absorbed by the atom in some manner, the electron might be able to move into another orbit having a higher energy.
  - The atom would then be in an EXCITED STATE.
- When this atom returns to its ground state it releases energy.



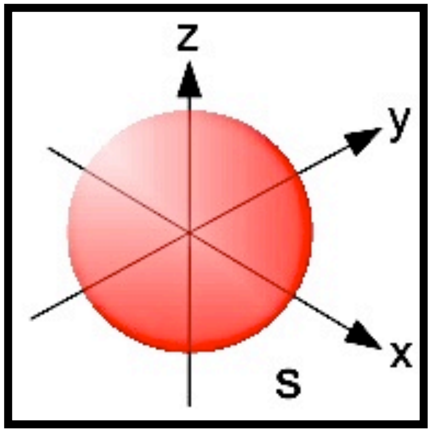
The Orbitals of the Electron Cloud

- Electrons within a shell do NOT follow the same path:
  - They occupy different sub-shells.
    - Within the ELECTRON CLOUD → Different ENERGY LEVELS or SHELLS
    - Within ENERGY LEVELS / SHELLS → Different SUB-SHELLS containing the electron's ORBITALS
      - Orbital = According to quantum mechanics, this is the regions were electrons are most probable to be found.
  - Each orbital can hold a maximum of two electrons.

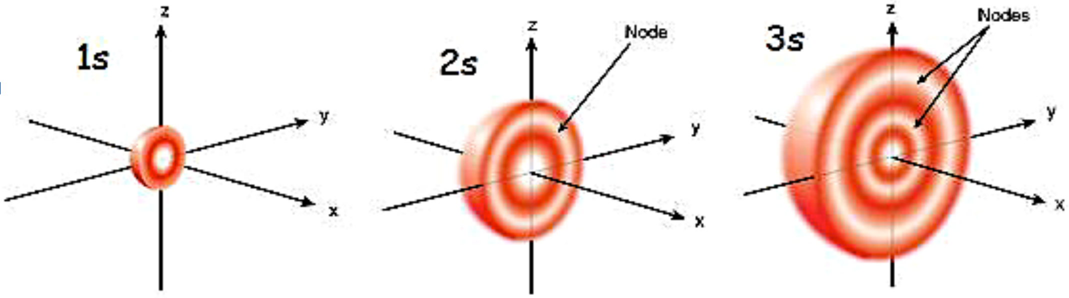


Electron Orbitals in Energy Level n = 1

- ORBITALS can hold only TWO electrons of opposite spins, electrons being repelled by nearby electrons, yet attracted to the protons in the nucleus

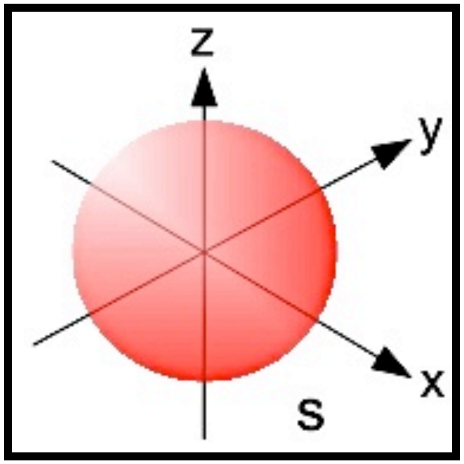


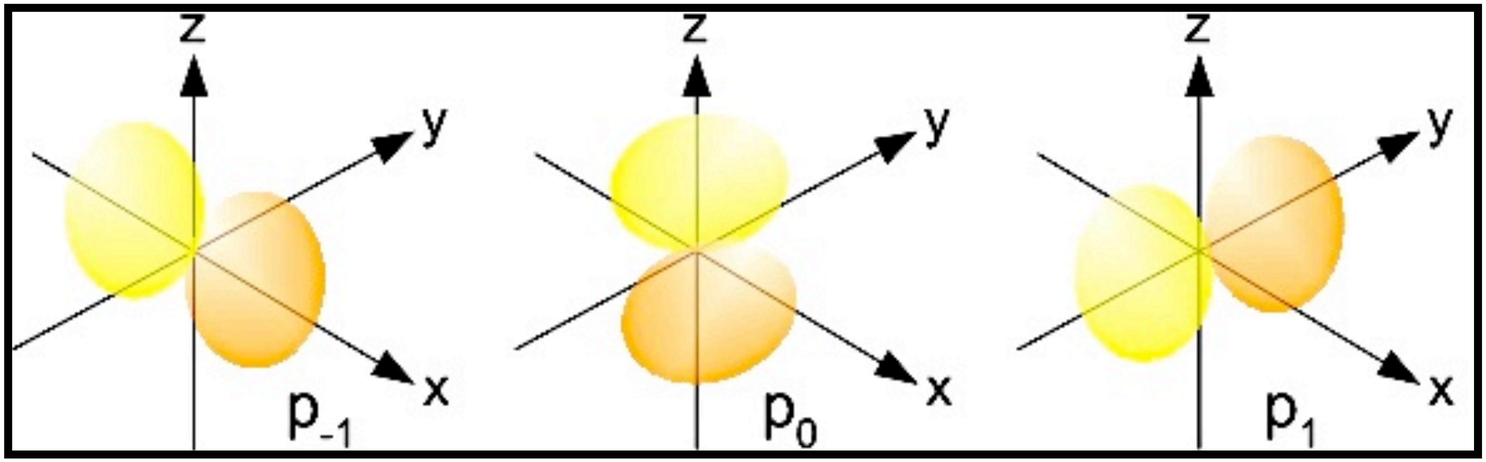
- Electron Orbitals: n = 1
  - In the 1st energy level, n = 1, there is only one sphere-shaped orbital: The 's' orbital.
    - An orbital exists if there are electrons in that region.
  - Additional s orbitals exist in certain atoms, each spherical region found further away from the nucleus.



Electron Orbitals in Energy Level n = 2

- Electron Orbitals: n = 2
  - In the 2nd energy level, n = 2, there is one sphere-shaped orbital (the 's' orbital) and three dumbbell-shaped orbitals (the 'p' orbital) that extend out of the nucleus at right angles, labeled the p_x, p_y, and p_z orbitals

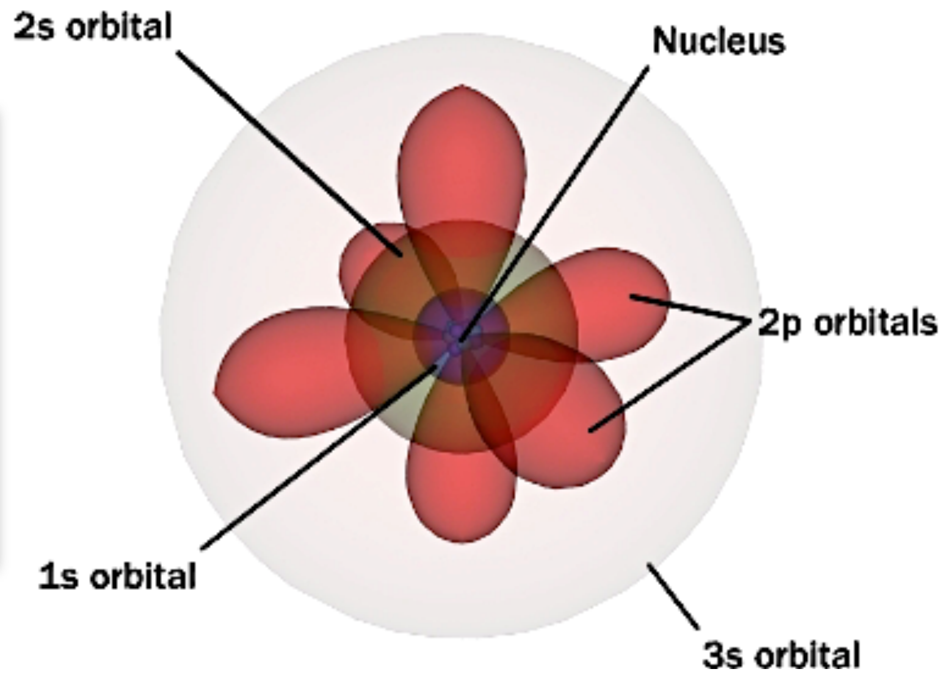




Electron Orbitals in Energy Level n = 2

The first energy level is made up of 1 spherical "s" orbital and can hold up to 2 electrons.

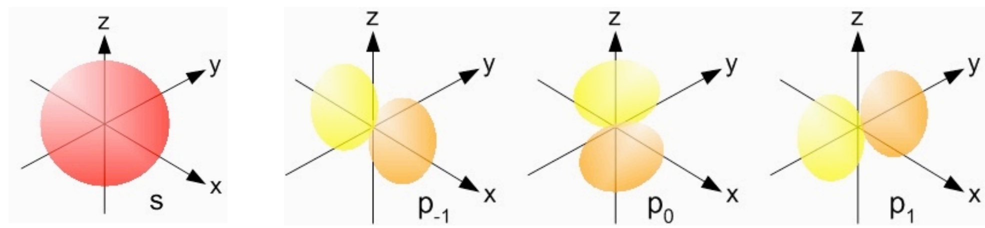
The second energy level is made up of 1 "s" orbital and up to 3 dumbbell shaped "p" orbitals and can hold up to 8 electrons.

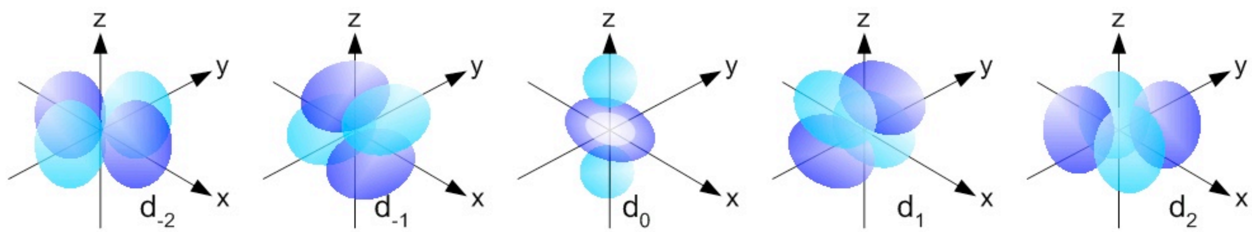


- Remember, as we move up in energy levels, the electrons are found <u>further</u> out from the nucleus.
  - For example, we find s and p orbitals in every energy level, they just get progressively larger in diameter as we move from lower to higher energy levels.

Electron Orbitals in Energy Level n = 3 & up

- **The larger energy levels have room for more electrons, which are held in additional orbitals such as the d orbitals and eventually the f orbitals.**
  - **The 3rd energy level, n = 3, can house up to one 's' orbital, three 'p' orbitals, and five 'd' orbitals**
  - **The 4th energy levels, n=4, can house up to one 's' orbital, three 'p' orbitals, five 'd' orbitals, and seven 'f' orbitals.**





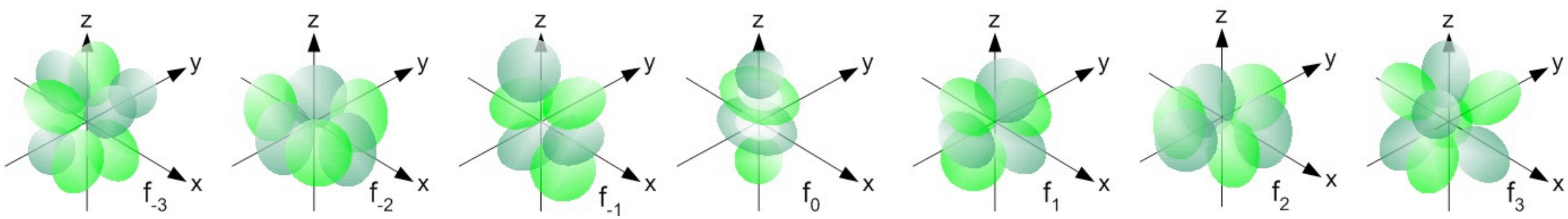
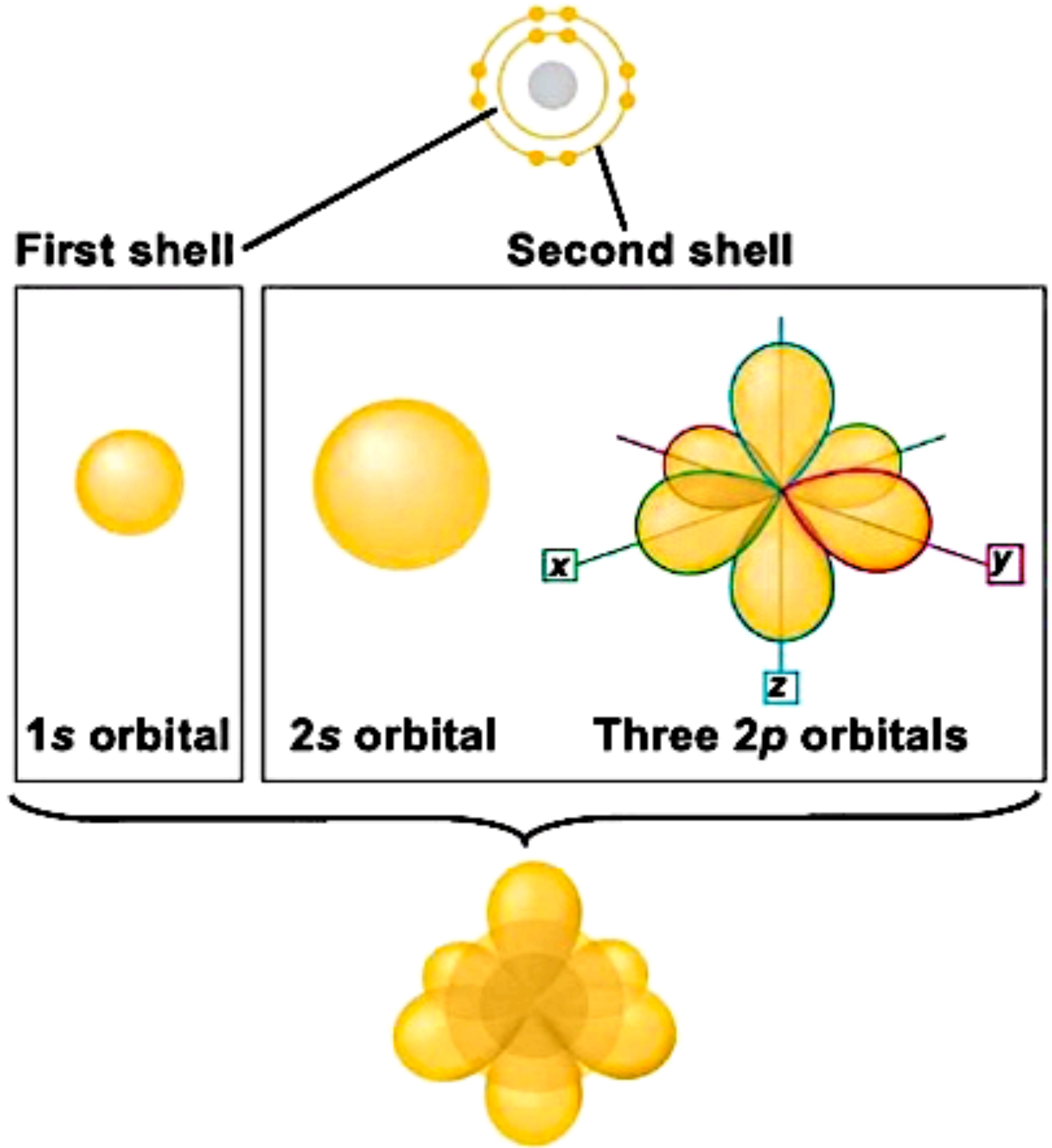


Fig. 2-10-4

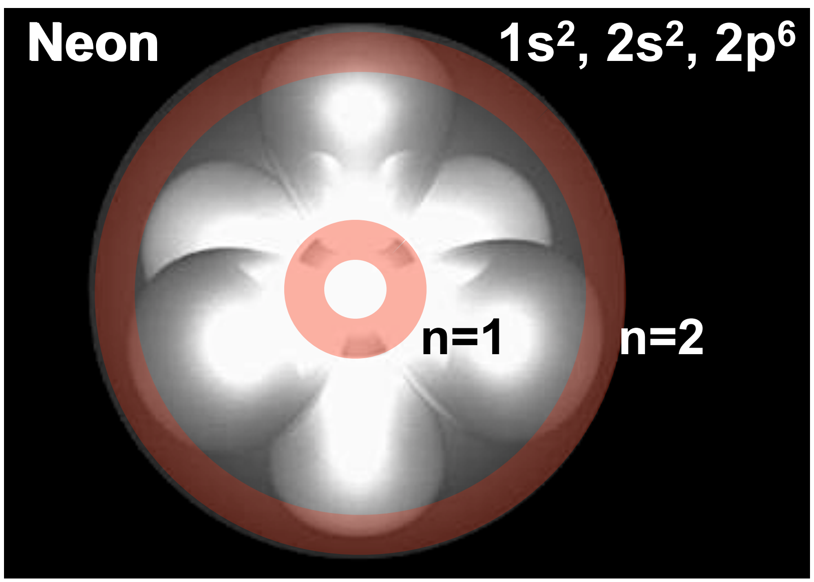
(a) Electron-distribution diagram

(b) Separate electron orbitals

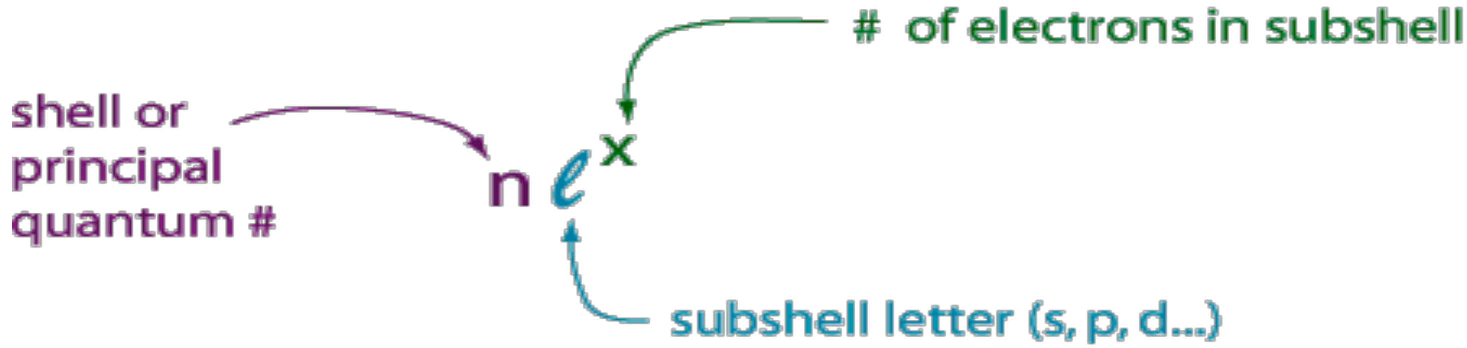
Neon, with two filled shells (10 electrons)



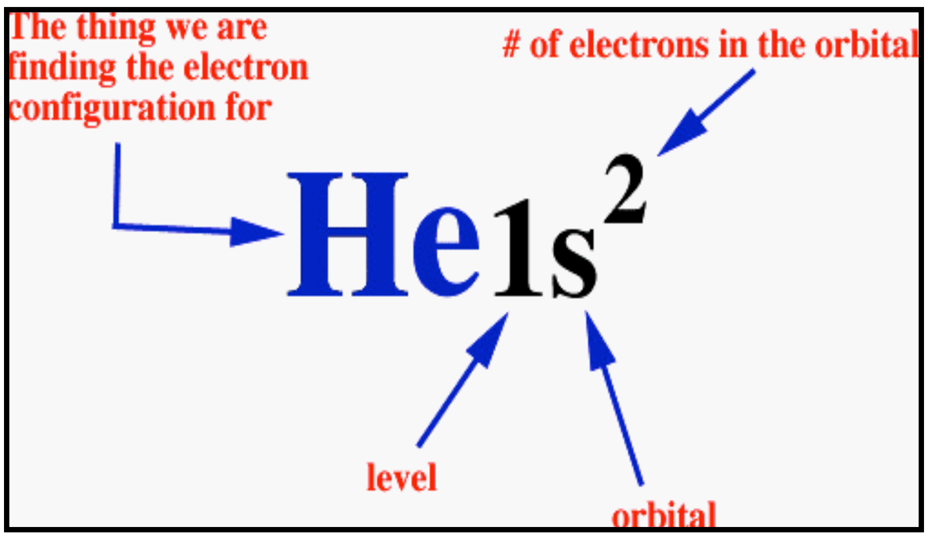
1s, 2s, and 2p orbitals



Electron Configuration Notation



EXAMPLE: Helium (He) → Atomic Number 2 → 2 protons & 2 electrons



EXAMPLE: What is the Electron Configuration of Lead (Pb)?

Pb = Atomic # 82. With 82 p, a neutral Pb has 82 e-

$1s^2$, $2s^2$, $2p^6$, $3s^2$, $3p^6$, $4s^2$, $3d^{10}$, $4p^6$, $5s^2$, $4d^{10}$, $5p^6$, $6s^2$, $4f^{14}$, $5d^{10}$, $6p^2$

What orbitals are occupied by the outermost electrons?

The outermost electrons ALWAYS occupy the outermost S and P orbitals

Sample Electron Configurations

- Notice how the electrons located in the outermost energy shell, what we call the valence electrons, are those occupying the s or the s and one, or more, p orbitals always

Which atoms have their outermost shell S and P electrons maximally filled?

S $1s^2, 2s^2, 2p^6, 3s^2, 3p^4$

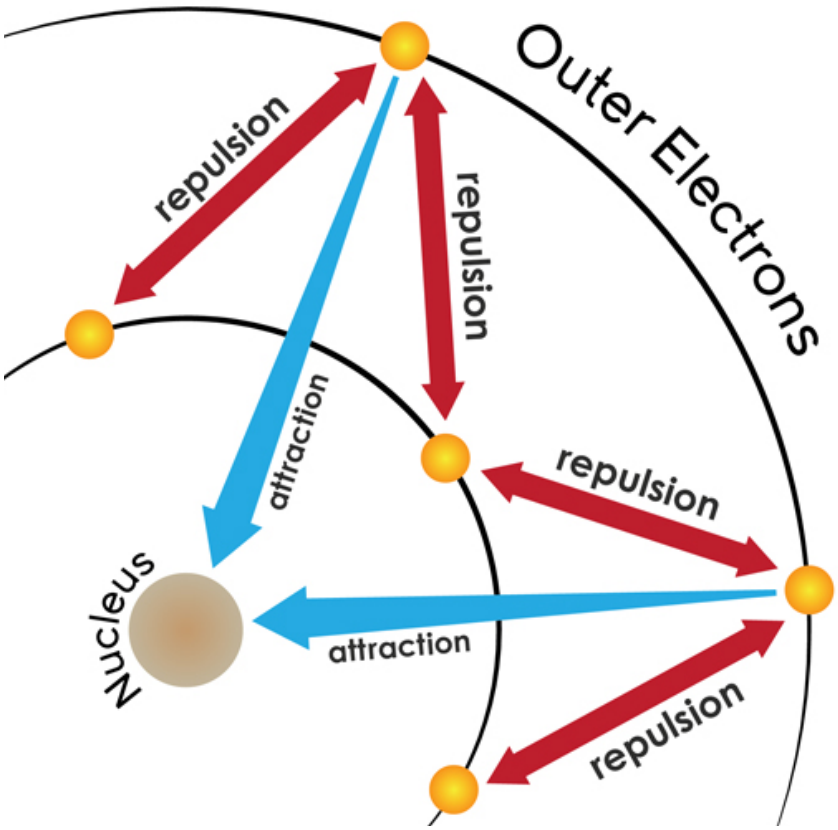
Rn $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^{14}, 5d^{10}, 6p^6$

Ac $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^{14}, 5d^{10}, 6p^6, 7s^2, 6d^1$

Hf $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^6, 5s^2, 4d^{10}, 5p^6, 6s^2, 4f^{14}, 5d^2$

Electron Orbitals & Valence Electrons

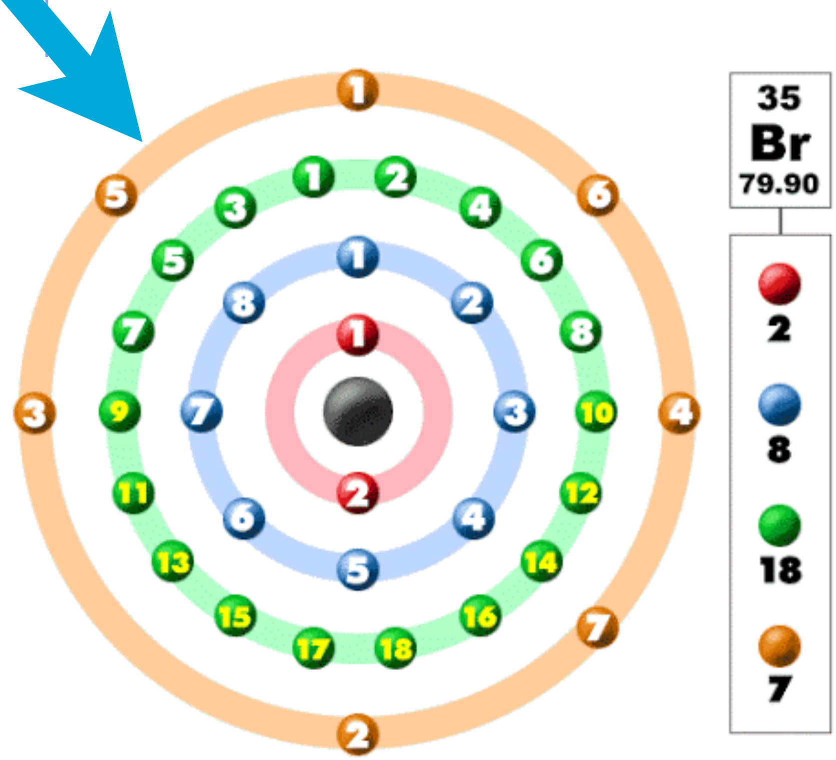
- Electrons are attracted to the protons in the nucleus, but still remain in their energy shells because of the electrostatic repulsions between the electrons in successive shells.

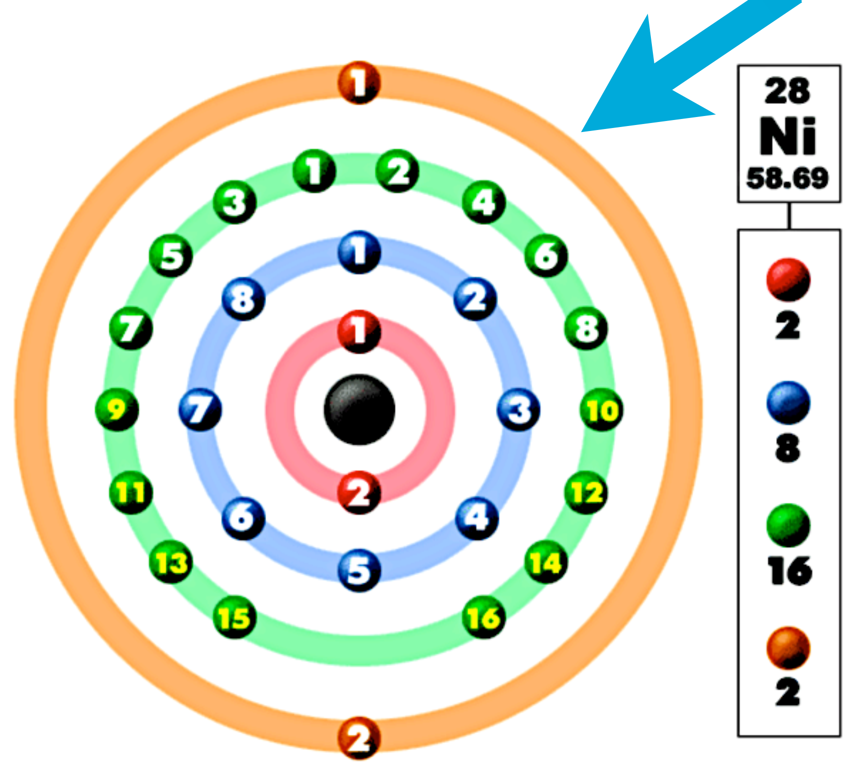


- Due to this electron-electron repulsion, electrons farthest from the nucleus, outermost shell electrons, are pushed father away and are thus held less tightly.
  - The outer electrons are SHIELDED by the inner energy level electrons from the attractive forces asserted by the positive protons in the nucleus.

Electron Orbitals & Valence Electrons

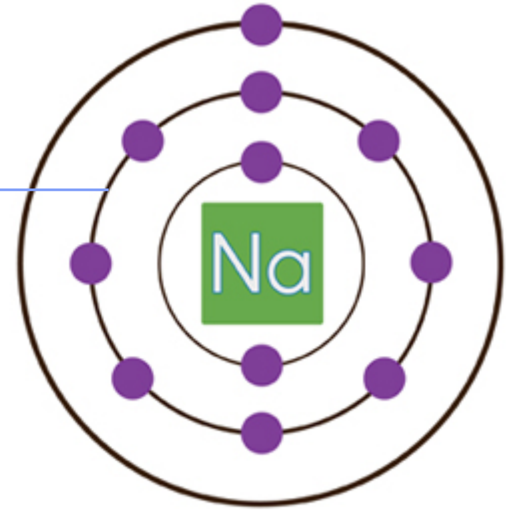
- The VALENCE SHELL is the outermost energy level of an atom.
  - The electrons that fill it, the VALENCE ELECTRONS, determine the chemical properties of the elements.

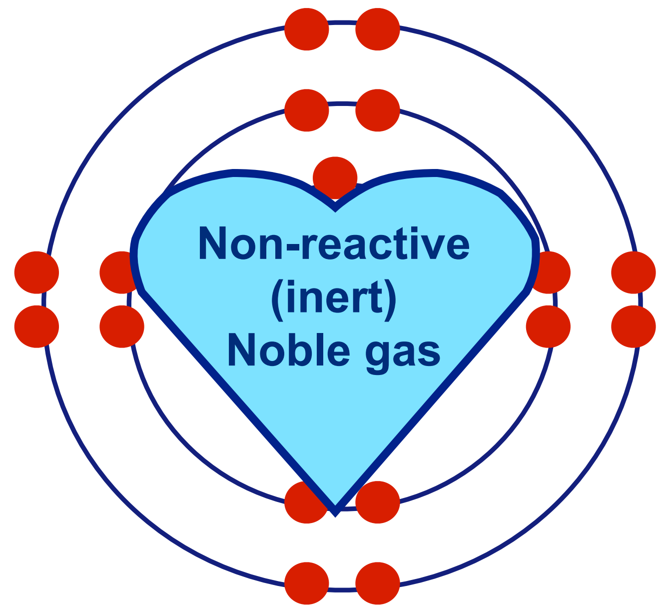




Bonding properties

- **Effect of electrons**
  - Being neutral (# e- = #p) does ***not*** mean the atom is most stable.
    - Atoms want the orbitals in their outermost energy shell (the Valence Shell) to be filled to be stable
    - Chemical behavior thus depends on the # of electron in the Valence Shell
      - For H, He, Li, Be etc.., this means having 2 electrons in their outer shell
      - For most atoms, though, this means having 8 electrons in the outer shell
        - Atoms will react with each other try to become more stable = to get a full valence shell
  - Electrons determine chemical behavior of atom *(how atoms behave around other atoms)*



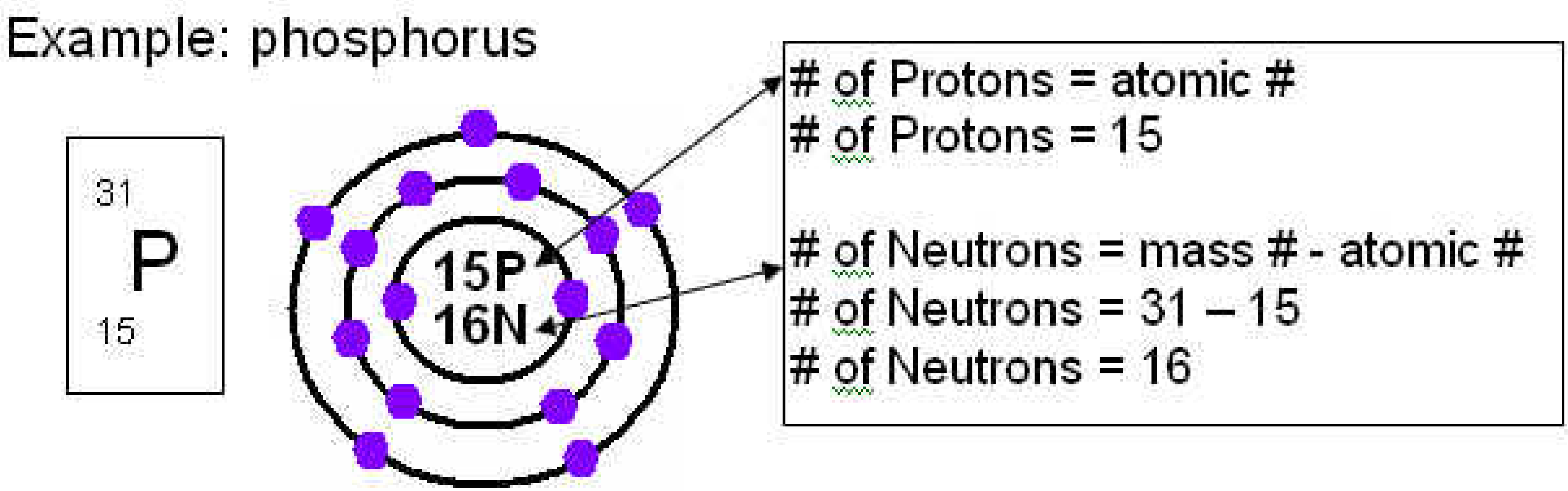


How do these atoms behave?

BOHR-RUTHERFORD DIAGRAM

- A BOHR-RUTHERFORD OR BOHR DIAGRAM is a way to visually represent the Electron Shells and the Electron Configuration of an atom.

  1. Draw the nucleus as a solid circle.
  2. Put the # of protons (Atomic #) in the nucleus with the # of neutrons (Mass # – Atomic #) under it.
  3. Place the # of electrons as dots on concentric circles around the nucleus representing the Electron Shells.



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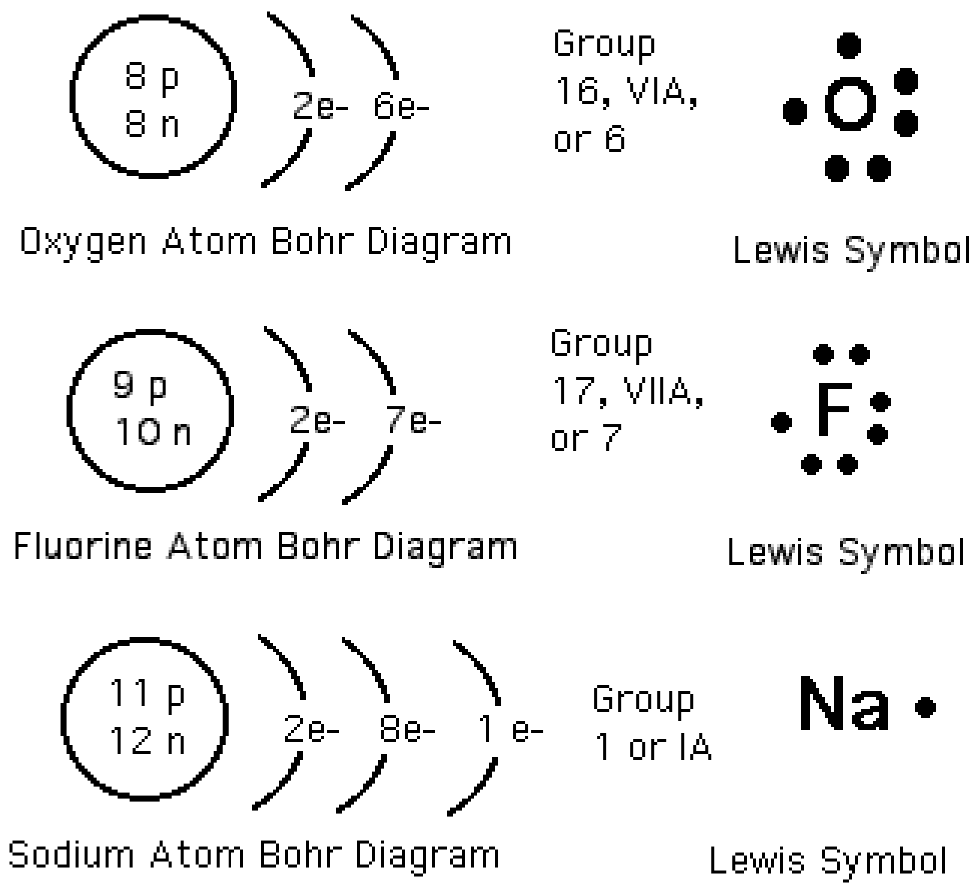
ELECTRON (LEWIS) - DOT STRUCTURES

- **Simplified Bohr diagrams which only consider electrons in outer energy levels (valence electrons) are called Lewis Symbols or Lewis Dot Structure.**



A Lewis Symbol consists of the element's symbol surrounded by "dots" to represent the number of electrons in the outer energy level as represented by a Bohr Diagram.

To draw the Lewis-Dot Structure, place one valence-electron-representing dot on each side of the atomic symbol first, doubling up the dots if you have more than four valence electrons ***(Exception - Helium has the two dots only on one side)***

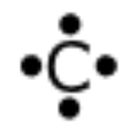
ELECTRON (LEWIS) - DOT STRUCTURES

- **Valence electrons are represented as dots placed on sides, top, bottom of the symbol for the element.**
  - **1, 2, 3 or 4 valence electrons are arranged around the element as a single dot on each side.**
    - **When there are more than 4 valence electrons, the dots are paired (no more than 2 per side).**

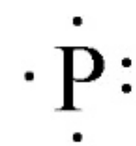
1 H							2 He
3 Li	4 Be	5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	81 Ti	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra						

- **Write the electron-dot structure of Hydrogen?**
  - **Symbol: H, Group 1A, so 1 valence e-**

- **Carbon?**
  - **Symbol: C, Group 4A, so 4 valence e-**

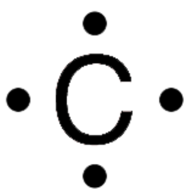


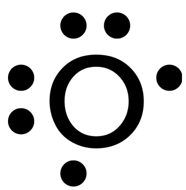
- **Phosphorous?**
  - **Symbol: P, Group 5A, so 5 valence e-**

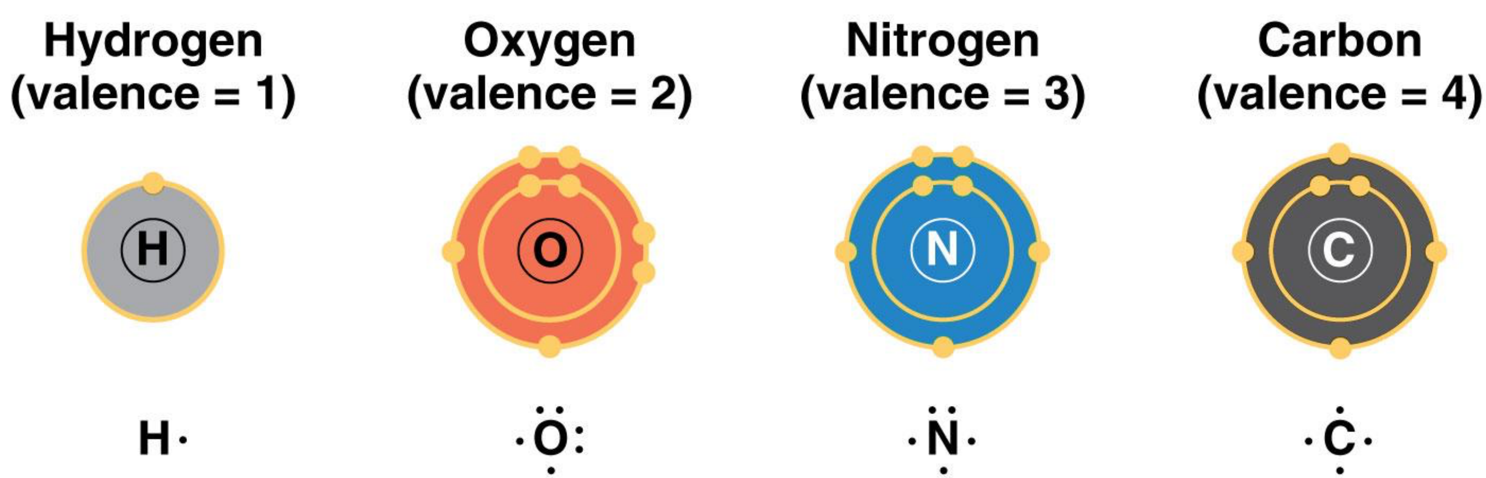


VALENCE & VALENCE ELECTRONS

- **The number of unpaired of electrons is called the atoms VALENCE.**
  - **This valence number is the number of covalent bonds that that atom is capable of forming.**
    - **EXAMPLE: Carbon, which has 6 electrons, has 4 valence electrons and a valence of 4 - it has 4 unpaired electrons.**
      - **The fact that C can form so many bonds explains why it forms the backbone of the organic molecules of life.**
    - **EXAMPLE: Oxygen which has 8 electrons, has 6 valence electrons and has a valence of 2 - it has 2 unpaired electrons.**

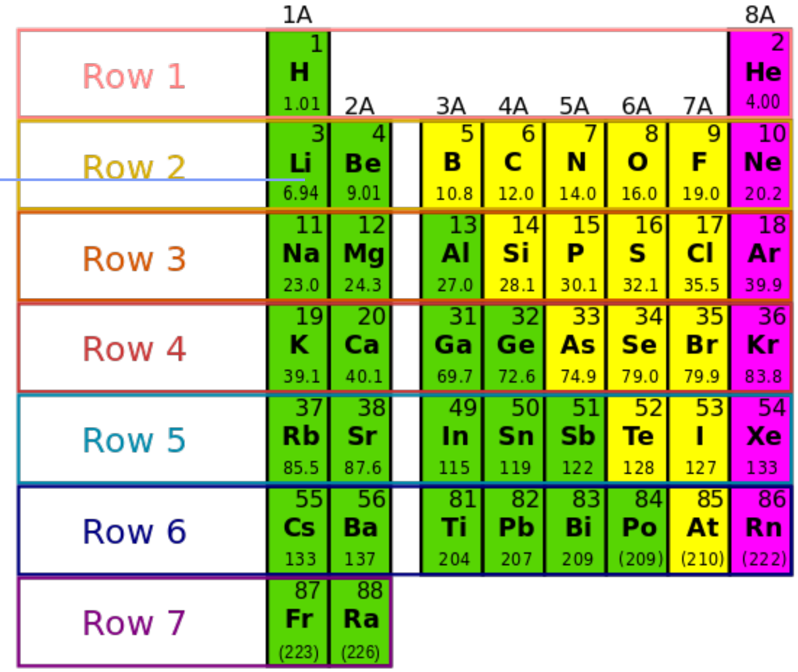






Bonding properties

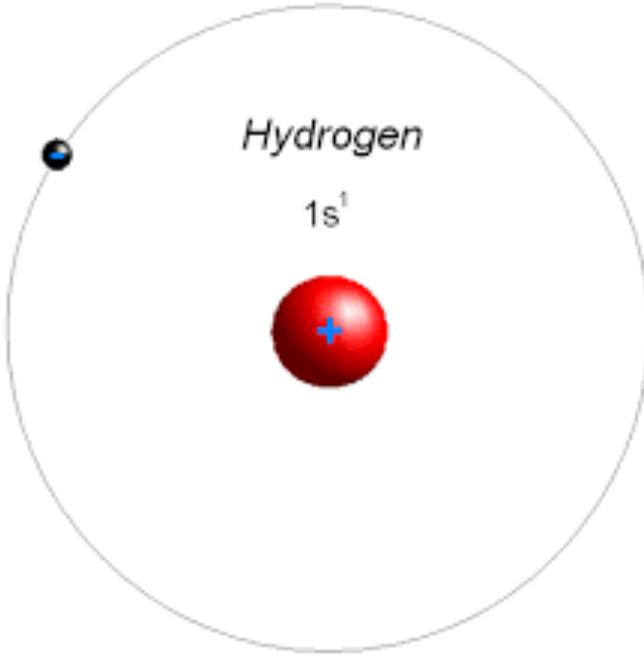
- **OCTET (& Duet) RULE:**
  - The noble (inert) gases (group 8A) already have a stable octet of 8 valence electrons due to their *filled outermost S orbital and three filled outermost P orbitals.*
  - Atoms of the other elements (including Groups 1A to 7A) form ions, compounds, or molecules together in an effort to be more stable, to have their outermost valence electron shell filled with the maximum amount of electrons.
    - For most this means having a total of 8 electrons in the valence shell *(for some it means having a total of 2 electrons in the valence shell)*

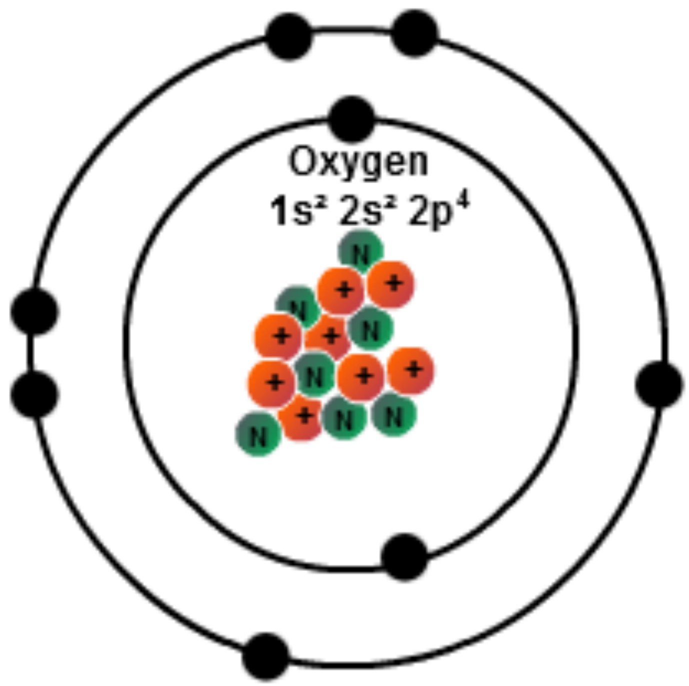


Group	Total Valence e-
All of 1A	1e-
All of 2A	2e-
All of 3A	3e-
All of 4A	4e-
All of 5A	5e-
All of 6A	6e-
All of 7A	7e-
All of 8A	8e-

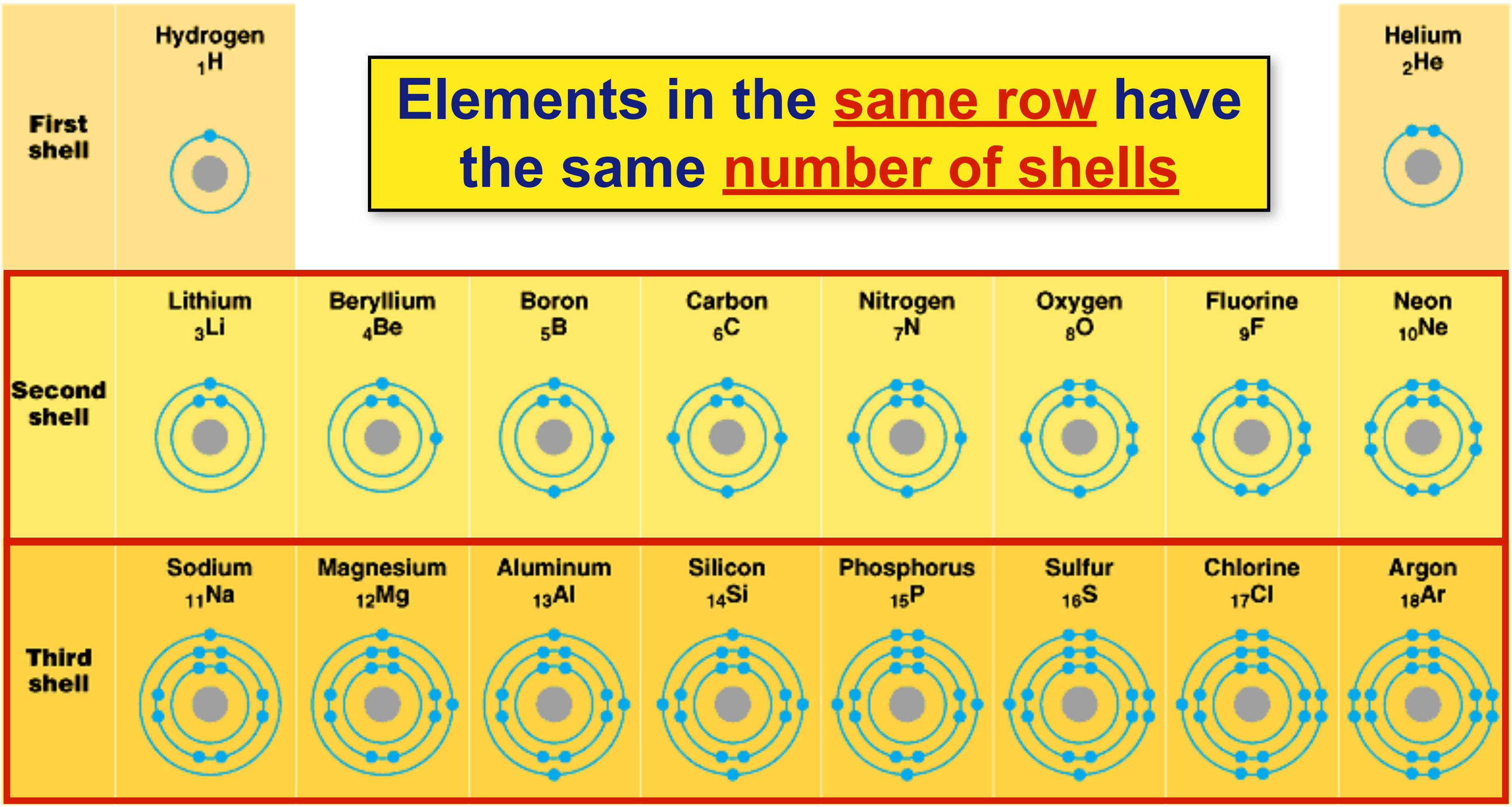
Bonding properties

- Ex: Hydrogen (H) and Helium (He) have 1 and 2 electrons, respectively, so their electrons only occupy the 1S orbital (the 1st energy shell cannot house any any P orbitals).
  - All orbitals can hold a maximum of 2 electrons.
    - H & He are stable with a duet of only 2 valence electrons and are said to follow the Duet rather than Octet Rule.
- Ex: Oxygen (O) has 8 total electrons, 6 of which are valence electrons.
    - O is stable with an octet of 8 valence electrons and is said to follow the Octet Rule.





Elements & their valence shells

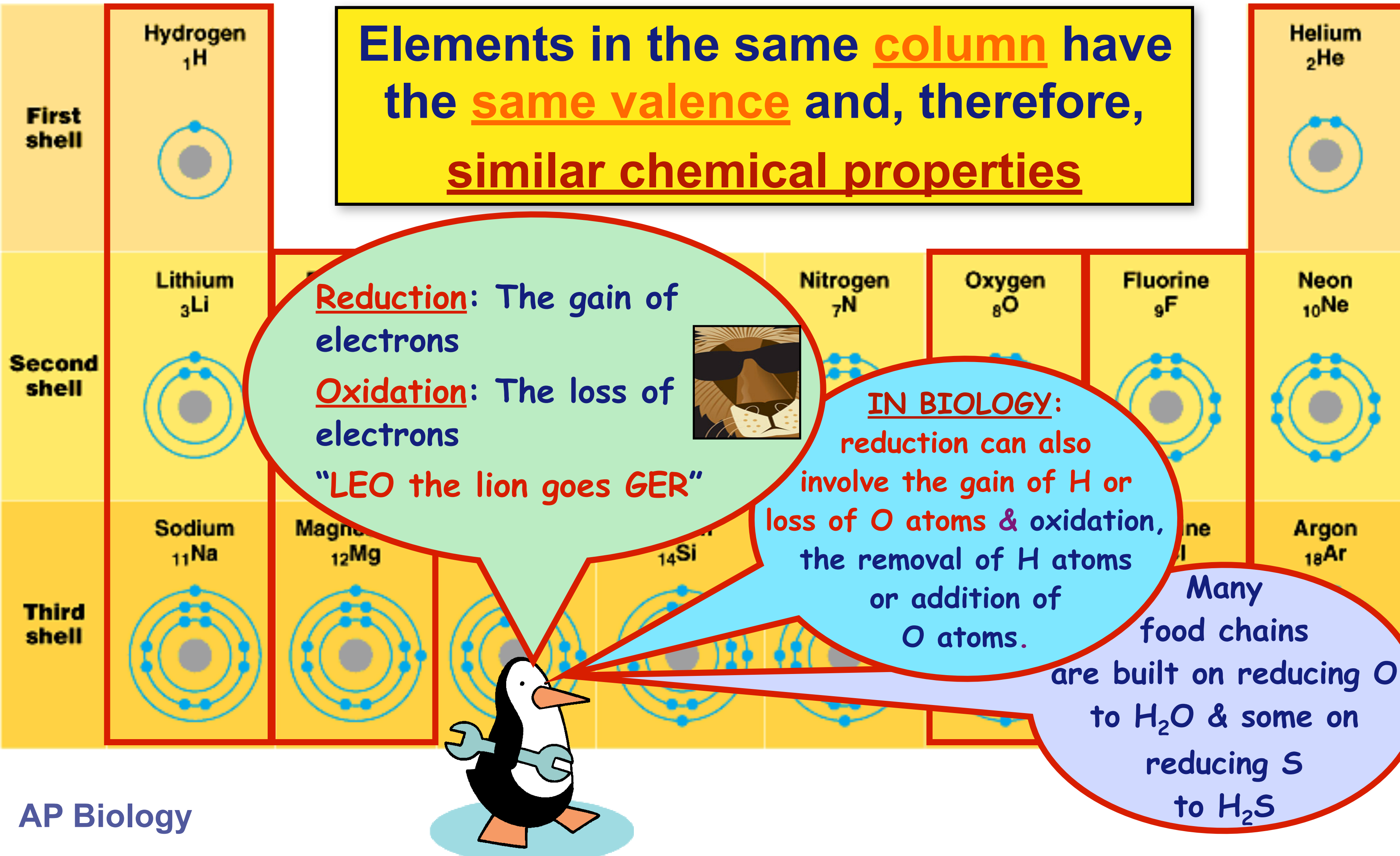


Elements in the same row have the same number of shells

Moving from left to right, each element has a sequential addition of protons *(each new proton attracting another electron)*.

Elements & their valence shells

Elements in the same column have the same valence and, therefore, similar chemical properties



Reduction: The gain of electrons

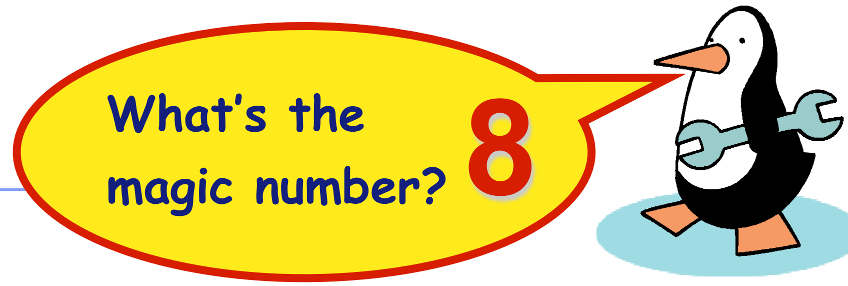
Oxidation: The loss of electrons

"LEO the lion goes GER"

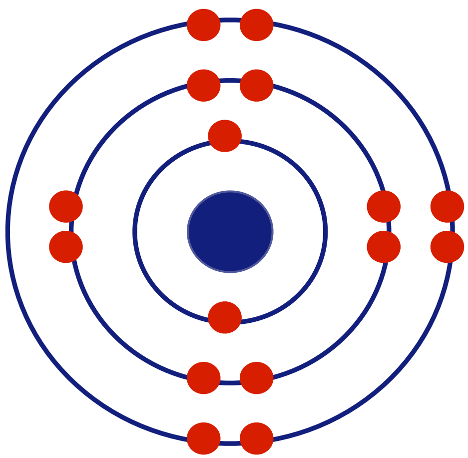
IN BIOLOGY: reduction can also involve the gain of H or loss of O atoms & oxidation, the removal of H atoms or addition of O atoms.

Many food chains are built on reducing O to H_2O & some on reducing S to H_2S

Bonding properties

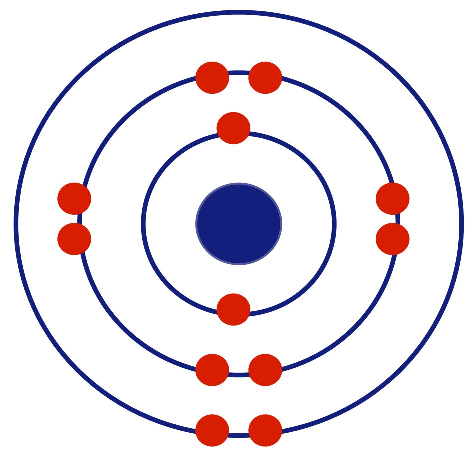


- **Effect of electrons**
  - The chemical behavior of an atom depends on number of electrons in its valence shell. To be stable, the orbitals in the outermost energy level/shell must be full. Remember the <u>OCTET (& DUET) Rules</u> - *atoms, with certain exceptions like Hydrogen & Helium, want 8 valence electrons to be stable.*



How does this atom behave?

With 6 valence e-'s, this atom, being a nonmetal, will either covalently bond or steal 2 additional electrons from another atom in order to complete its valence shell (the third Electron Shell)



How does this atom behave?

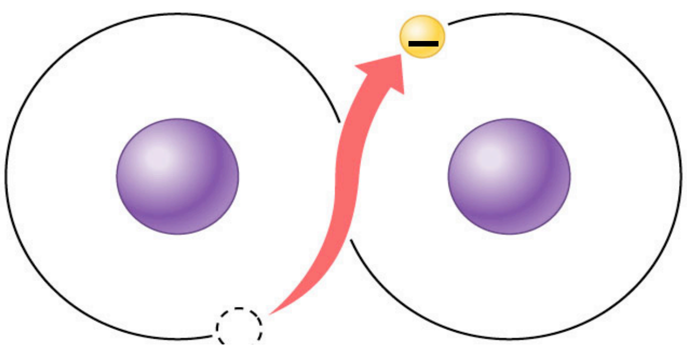
With 2 valence e-'s, and being a metal, this atom will lose both valence e-'s becoming a positively charged ion (cation) with a 2+ charge that now has a full valence shell (The second Electron Shell)

Chemical Reactivity

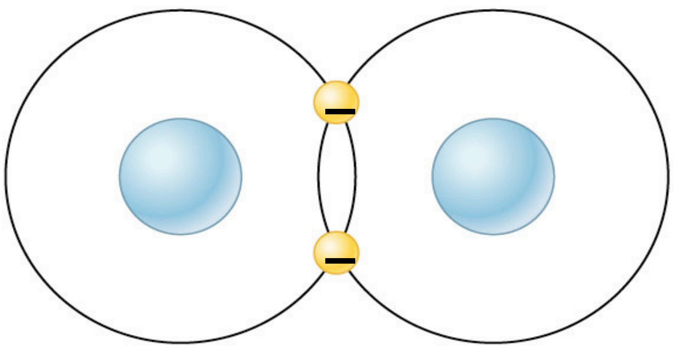
- To become more stable, atoms will react to...
  - complete a partially filled valence shell or
  - empty a partially filled valence shell

This tendency drives chemical reactions…

and creates chemical bonds (each bond forms between 2 atoms)

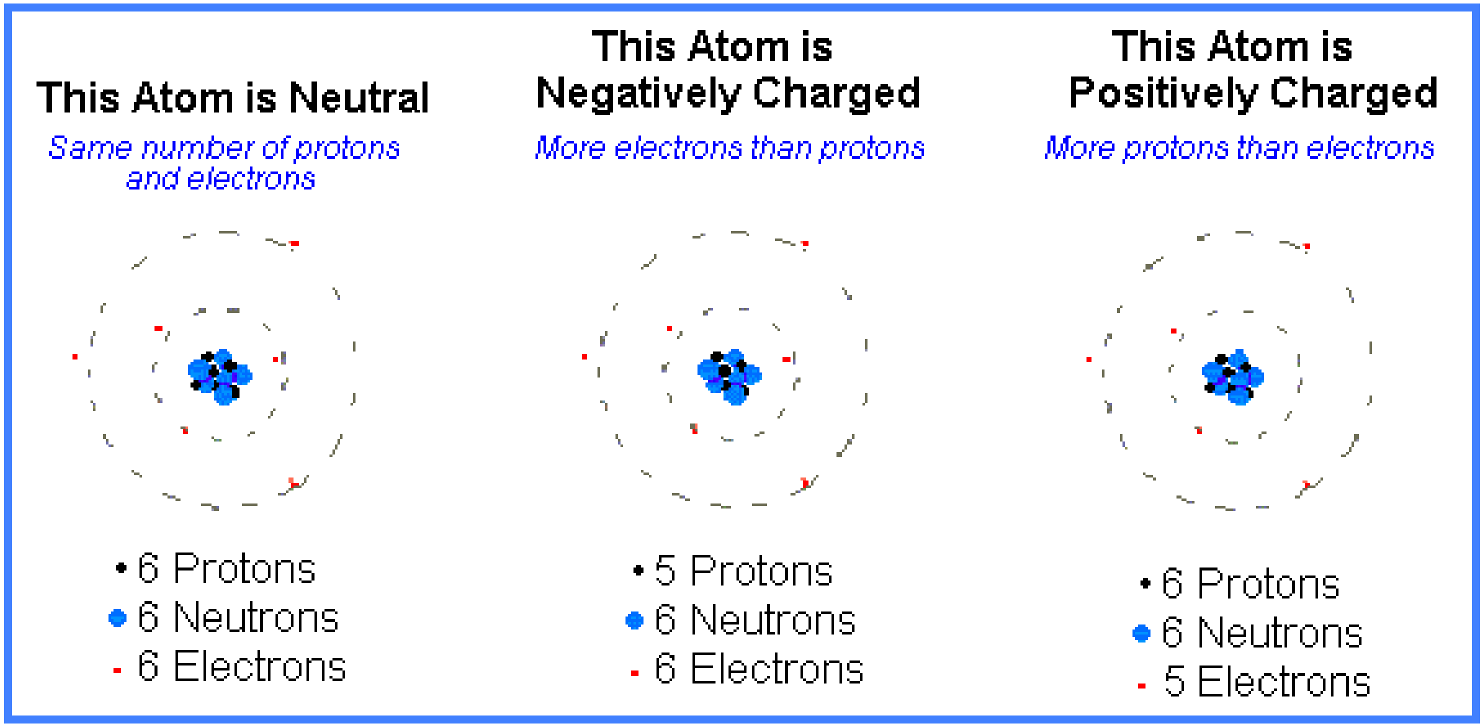


Ionic bonds form when one atom steals electrons from another atom. Both now become oppositely charged ions, which are now attracted each other.



Covalent bonds form when two atoms share electrons in order for each to complete their valence shells.

Ions are charged atoms or molecules



All charges in a neutral atom add up to <u>ZERO</u>! (#p = #e-)

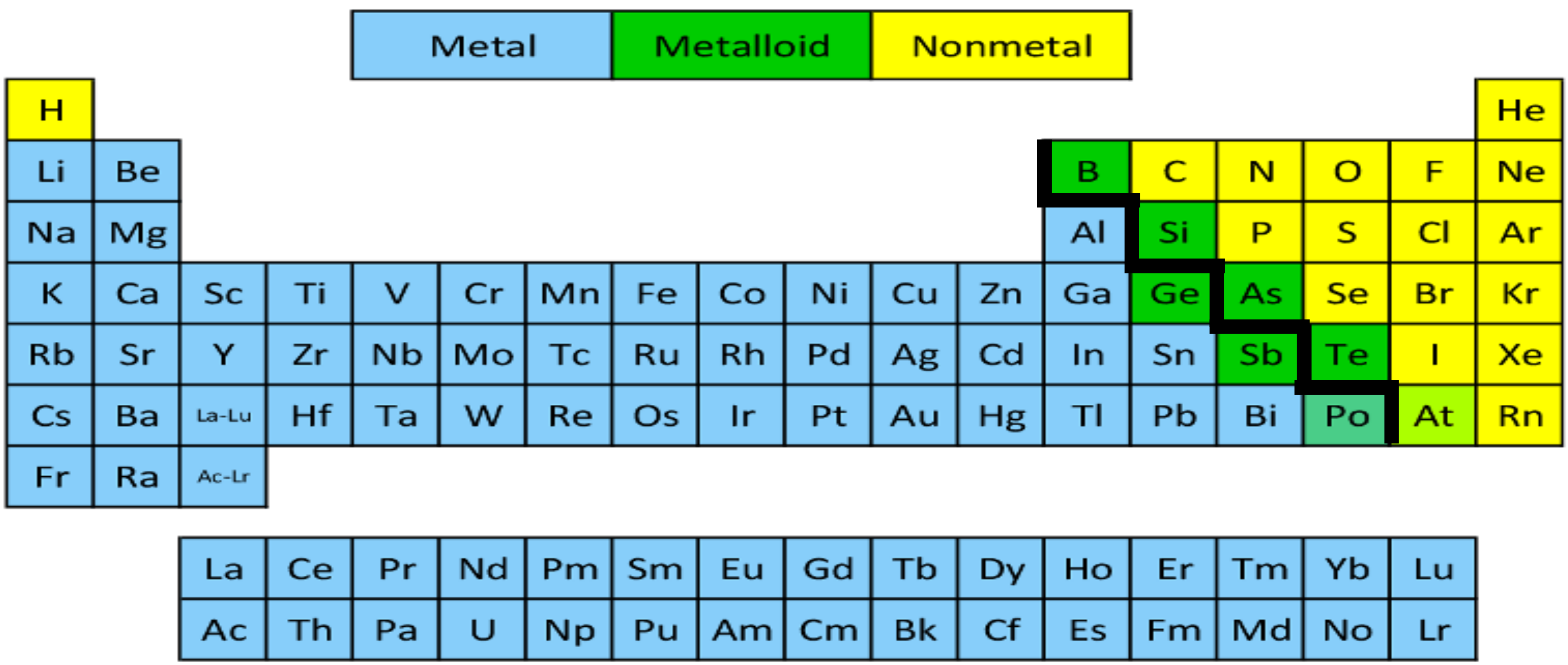
Ions form when an atom gains or loses electrons. (#p ≠ #e-)

<u>Cations</u> = positively charged ions

<u>Anions</u> = negatively charged ions

The Periodic Table of Elements

1. Except for hydrogen (H) on the top left, elements on the left-hand side of the zigzag line, which runs from Boron (B) to astatine (At), are metals. (Blue)
2. All elements to the right of the zigzag line and Hydrogen (H) on the top left are nonmetals. (Yellow)
3. Seven of the nine elements that touch the zigzag line are semimetals. (Green)



The Periodic Table of Elements

EXERCISE: **Are the following elements metals, nonmetals, or semimetals (metalloids) ?**

Ca (Calcium) ?
H (Hydrogen) ?
Si (Silicone) ?
Fe (Iron) ?
O (Oxygen) ?

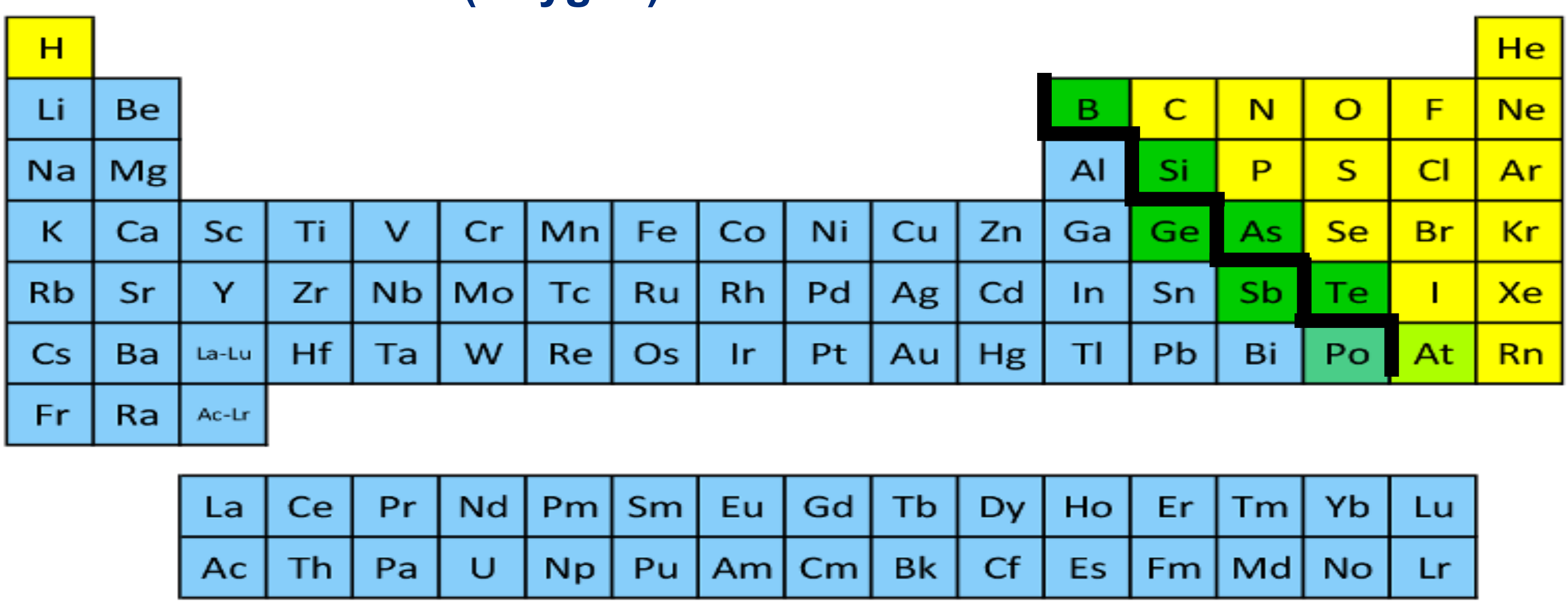
H																	He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La-Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac-Lr															

La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

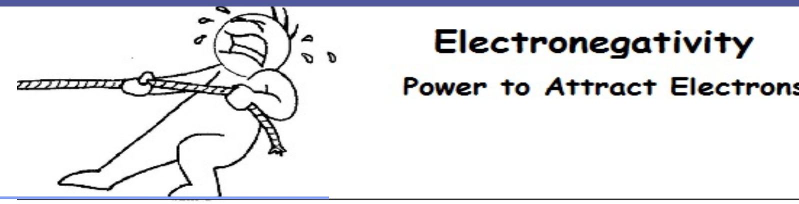
The Periodic Table of Elements

EXERCISE: **Are the following elements metals, nonmetals, or semimetals (metalloids) ?**

Ca (Calcium) ?	**Metal**
H (Hydrogen) ?	**Nonmetal**
Si (Silicone) ?	**Metalloid**
Fe (Iron) ?	**Metal**
O (Oxygen) ?	**Nonmetal**



Electronegativity



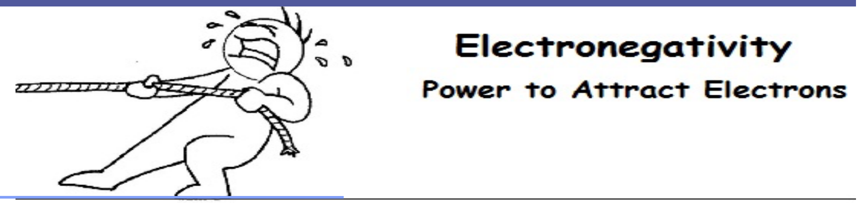
- <u>Electronegativity</u> **is a measure of the tendency of an atom to attract a bonding pair of electrons** *(a measure of "how strong a nucleus pulls on electrons")*
  - ***Electronegativity ranges in value from 0.7 to 4.0***
    - **Generally, <u>metals</u> tend to be the <u>LEAST</u> electronegative while <u>nonmetals</u> tend to be the <u>MOST</u> electronegative**
      - **Atoms may lose or gain electrons in order to ensure their**

Electronegativity values of the elements (Pauling scale)

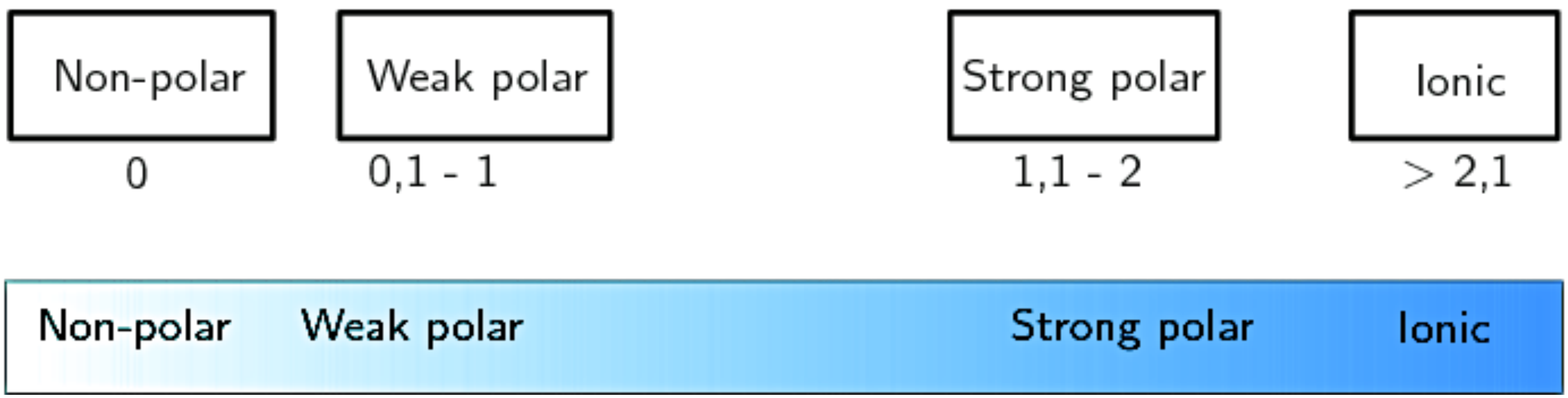
H 2.1																	He
[illegible]	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
Na 0.9	Mg 1.2											Al 1.5	[illegible] 1.8	[illegible]	[illegible] 2.5	Cl 3.0	Ar
[illegible] 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu [illegible]	[illegible]	[illegible] 1.6	[illegible] 1.8	[illegible] 2.0	Se 2.4	Br 2.8	Kr 3.0
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru [illegible]	[illegible]	[illegible] 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe 2.6
Cs 0.7	Ba 0.9	La 1.1	Hf [illegible]	Ta [illegible]	[illegible] 1.7	[illegible] 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Ti 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn 2.4
Fr 0.7	Ra [illegible]	[illegible] 1.1															

Ce 1.1	Pr 1.1	Nd 1.1	Pm 1.1	Sm 1.1	Eu 1.1	Gd 1.1	Tb 1.1	Dy 1.1	Ho 1.1	Er 1.1	Tm 1.1	Yb 1.1	Lu 1.2
Th 1.3	Pa 1.5	U 1.7	Np 1.3	Pu 1.3	Am 1.3	Cm 1.3	Bk 1.3	Cf 1.3	Es 1.3	Fm 1.3	Md 1.3	No 1.3	Lr

Electronegativity



- **<u>When two atoms near each other, the difference in their Electronegativity will determine what kind of bond results.</u>**
  - **Large differences lead to one atom stealing valence electrons from another, leading to ion formation and, thus, <u>ionic bonding</u>**
  - **Small difference lead to two atoms sharing the electron more equally as part of a <u>non-polar covalent bond</u>**
  - **Medium differences lead to two atoms sharing electrons more unequally as part of a <u>polar covalent bond</u>**



Positive Ions

Recall that the # of valence electrons in groups 1A to 8A is equal to the group number.

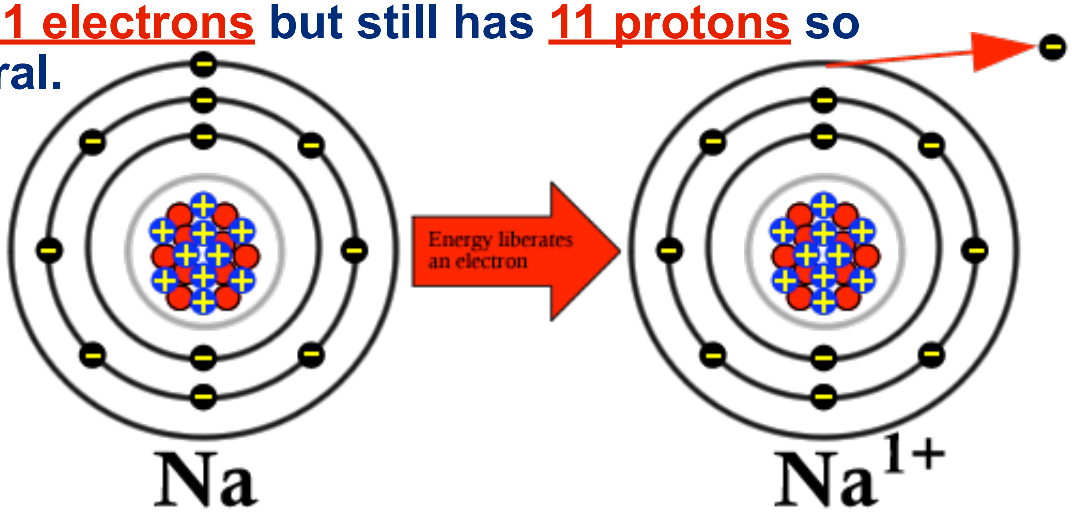
- **Metals**
  - **Low Electronegativity** (= how strongly an atom's nucleus pulls on electrons)
    - **LOSE** their valence electrons to much more electronegative nonmetal atoms, attaining a noble gas electron configuration, forming **CATIONS** *(positively charged atoms)*

Example: Sodium (Na) is in Group 1A so it has 1 valence e- instead of 8. It looses its valence e- to a much more electronegative atom (a nonmetal) and becomes a positive ion.

Na now has 10 instead of 11 electrons but still has 11 protons so the atom is no longer neutral.

It carries an electrical charge, or ionic charge, of 1+.

Ionic charge is written in the upper right-hand corner of the element's symbol



Has 1 valence e- in the 3s orbital

Now, has 8 valence e-'s in the 2s & all three 2p orbitals

Positive Ions

Electron configuration of neutral Na: Na $1s^2$, $2s^2$, $2p^6$, $3s^1$

With only 1 electron in its outer most valence shell (Energy Level n=3), Na is unstable. To be stable, sodium obeys the Octet Rule and wants all s and p orbitals filled in it's outermost, valence shell.

Na can either gain 7 extra e-'s or just loose 1. Because it is low in electronegativity, it will loose the 1 valence e- easily to another more electronegative atom (a nonmetal). This deletes the third Electron Shell entirely, leaving sodium with a new electron configuration that is similar to that of the stable Noble Gas Neon with 8 valence e-'s in its valence shell:

Na $1s^2$, $2s^2$, $2p^6$, ~~$3s^1$~~ → Na $1s^2$, $2s^2$, $2p^6$

Ne $1s^2$, $2s^2$, $2p^6$

Positive Ions

Na $1s^2, 2s^2, 2p^6, 3s^1$ → Na $1s^2, 2s^2, 2p^6$

Note that the number of protons did NOT change, but the number of electrons did.

Total p = 11
Total e- in neutral Na = 11
New Total e- after loosing 1 valence e- = 11 - 1 = 10

Net Charge on new Na Atom: (+11) + (-10) = +1

The neutral sodium atom now has a charge imbalance and is no longer neutral, but has become an ion, or a cation to be more specific.

The symbol for the ion of sodium is Na^{1+} or Na^{+}

Negative Ions

- NON-Metals
  - High Electronegativity
    - GAIN one or more valence electron(s) from much less electronegative metal atoms in order to attain a noble gas forming negatively charged ions or ANIONS.
      (negatively charged atoms or ions)

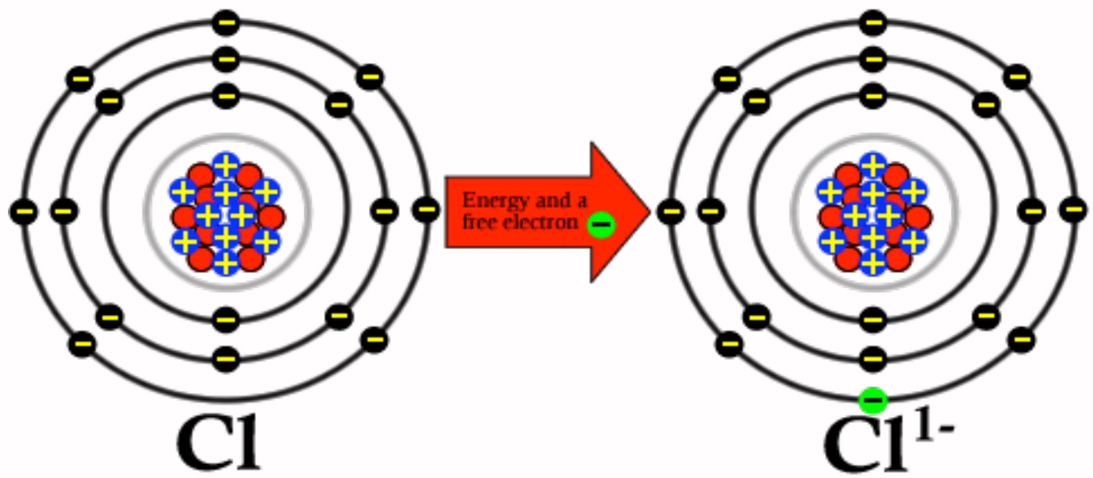
- EXAMPLE: Cl = Belongs to Group 7A so has 7 valence electrons

When it gains an extra outer e^-, it will achieve noble gas arrangement like Argon (Ar) of Group 8A becoming a negatively charged ion.

Cl now has 18 instead of 17 electrons but still has 17 protons so the atom is no longer neutral.

It carries an electrical charge, or ionic charge, of 1-.

Ionic charge is written in the upper right-hand corner of the element's symbol



Has 7 valence e-'s in the 3s & three 3p orbitals

Now, has 8 valence e-'s in the 3s & three 3p orbitals

Negative Ions

Electron configuration of neutral Cl: Cl $1s^2, 2s^2, 2p^6, 3s^2, 3p^5$

Cl is unstable because it has only 7 e- in its outermost valence shell, Energy Level n=3. To be stable, chlorine obeys the Octet Rule and wants all s and p orbitals filled in it's outermost, valence shell.

Cl can either gain 1 extra e- or loose 7 e-'s to obtain a fully filled valence shell. It will gain the 1 valence e-, stealing the electron from a much less electronegative atom (a metal), leaving chlorine with a new electron configuration that is similar to that of the stable Noble Gas Atgon with 8 valence e-'s in its valence shell:

Cl $1s^2, 2s^2, 2p^6, 3s^2, 3p^5$ → Cl $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

Ar $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

Negative Ions

Cl $1s^2, 2s^2, 2p^6, 3s^2, 3p^5$ → Cl $1s^2, 2s^2, 2p^6, 3s^2, 3p^6$

Note that the number of protons did NOT change but the number of electrons did.

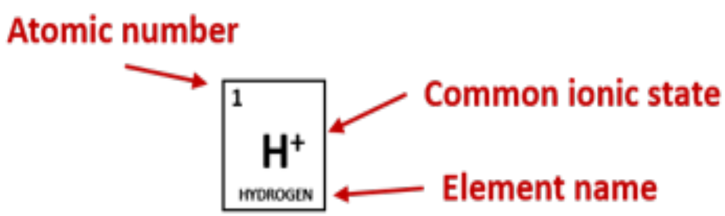
Total p = 17
Total e- in neutral Cl = 17
New Total e- after gaining 1 valence e- = 17 + 1 = 18

Net Charge on new Cl Atom: (+17) + (-18) = -1

The neutral chlorine atom now has a charge imbalance and is no longer neutral but has become an ion or a anion to be more specific.

The symbol for the ion of chlorine is Cl^{1-} or Cl^{-}

Common Ionic States of Elements



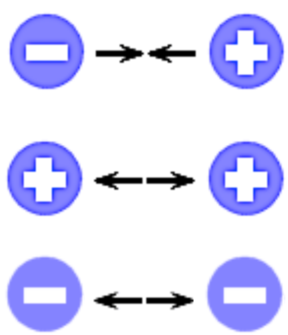
Many of the transition metals can take multiple charges.

+1	+2											+3		-3	-2	-1	
1 H^{+} HYDROGEN																	2 He HELIUM
3 Li^{+} LITHIUM	4 Be^{2+} BERYLLIUM											5 B BORON	6 C CARBON	7 N^{3-} NITROGEN	8 O^{2-} OXYGEN	9 F^{-} FLUORINE	10 Ne NEON
11 Na^{+} SODIUM	12 Mg^{2+} MAGNESIUM											13 Al^{3+} ALUMINIUM	14 Si SILICON	15 P^{3-} PHOSPHORUS	16 S^{2-} SULFUR	17 Cl^{-} CHLORINE	18 Ar ARGON
19 K^{+} POTASSIUM	20 Ca^{2+} CALCIUM	21 Sc^{3+} SCANDIUM	22 Ti^{3+} / Ti^{4+} TITANIUM	23 V^{3+} / V^{5+} VANADIUM	24 Cr^{2+} / Cr^{3+} CHROMIUM	25 Mn^{2+} / Mn^{4+} MANGANESE	26 Fe^{2+} / Fe^{3+} IRON	27 Co^{2+} / Co^{3+} COBALT	28 Ni^{2+} / Ni^{3+} NICKEL	29 Cu^{+} / Cu^{2+} COPPER	30 Zn^{2+} ZINC	31 Ga^{3+} GALLIUM	32 Ge^{4+} GERMANIUM	33 As^{3-} ARSENIC	34 Se^{2-} SELENIUM	35 Br^{-} BROMINE	36 Kr KRYPTON
37 Rb^{+} RUBIDIUM	38 Sr^{2+} STRONTIUM	39 Y^{3+} YTTRIUM	40 Zr^{4+} ZIRCONIUM	41 Nb^{3+} / Nb^{5+} NIOBIUM	42 Mo^{6+} MOLYBDENUM	43 Tc^{7+} TECHNETIUM	44 Ru^{3+} / Ru^{4+} RUTHENIUM	45 Rh^{3+} RHODIUM	46 Pd^{2+} / Pd^{4+} PALLADIUM	47 Ag^{+} SILVER	48 Cd^{2+} CADMIUM	49 In^{3+} INDIUM	50 Sn^{2+} / Sn^{4+} TIN	51 Sb^{3+} / Sb^{5+} ANTIMONY	52 Te^{2-} TELLURIUM	53 I^{-} IODINE	54 Xe XENON
55 Cs^{+} CESIUM	56 Ba^{2+} BARIUM	71 Lu^{3+} LUTETIUM	72 Hf^{4+} HAFNIUM	73 Ta^{5+} TANTALUM	74 W^{6+} TUNGSTEN	75 Re^{7+} RHENIUM	76 Os^{4+} OSMIUM	77 Ir^{4+} IRIDIUM	78 Pt^{2+} / Pt^{4+} PLATINUM	79 Au^{+} / Au^{3+} GOLD	80 Hg_2^{2+} / Hg^{2+} MERCURY	81 Tl^{+} / Tl^{3+} THALLIUM	82 Pb^{2+} / Pb^{4+} LEAD	83 Bi^{3+} / Bi^{5+} BISMUTH	84 Po^{2+} / Po^{4+} POLONIUM	85 At^{-} ASTATINE	86 Rn RADON
87 Fr^{+} FRANCIUM	88 Ra^{2+} RADIUM	103 Lr^{3+} LAWRENCIUM															

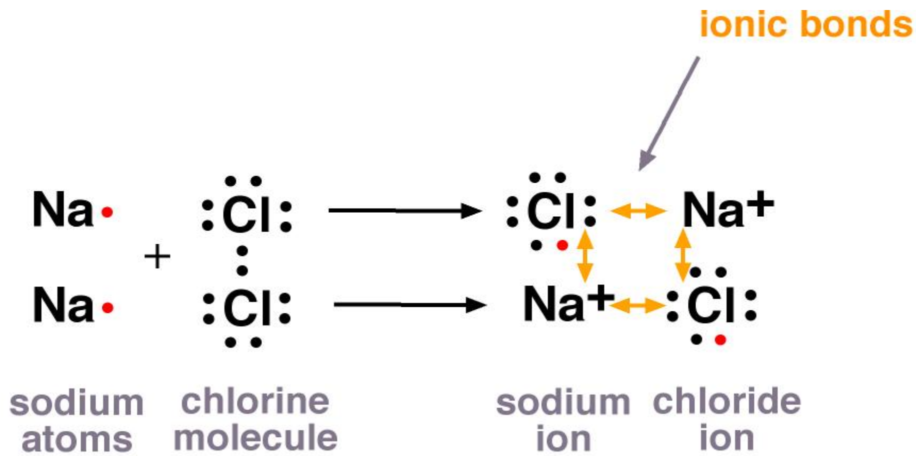
57 La^{3+} LANTHANUM	58 Ce^{3+} CERIUM	59 Pr^{3+} PRASEODYMIUM	60 Nd^{3+} NEODYMIUM	61 Pm^{3+} PROMETHIUM	62 Sm^{2+} / Sm^{3+} SAMARIUM	63 Eu^{2+} / Eu^{3+} EUROPIUM	64 Gd^{3+} GADOLINIUM	65 Tb^{3+} TERBIUM	66 Dy^{3+} DYSPROSIUM	67 Ho^{3+} HOLMIUM	68 Er^{3+} ERBIUM	69 Tm^{3+} THULIUM	70 Yb^{3+} YTTERBIUM
89 Ac^{3+} ACTINIUM	90 Th^{4+} THORIUM	91 Pa^{4+} / Pa^{5+} PROTACTINIUM	92 U^{4+} / U^{6+} URANIUM	93 Np^{5+} NEPTUNIUM	94 Pu^{4+} / Pu^{4+} PLUTONIUM	95 Am^{3+} / Am^{4+} AMERICIUM	96 Cm^{3+} CURIUM	97 Bk^{3+} BERKELIUM	98 Cf^{3+} CALIFORNIUM	99 Es^{3+} EINSTEINIUM	100 Fm^{3+} FERMIUM	101 Md^{2+} / Md^{3+} MENDELEVIUM	102 No^{2+} / No^{3+} NOBELIUM

Opposites Attract

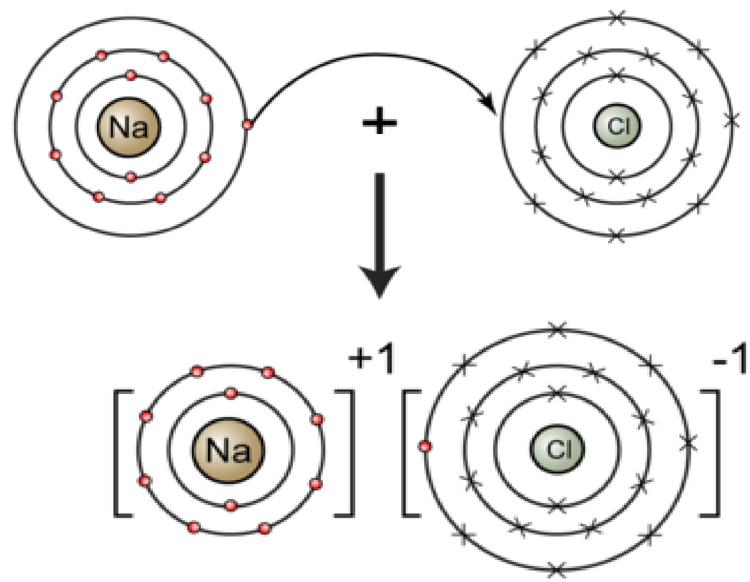
Oppositely charged particles will attract each other.



- For <u>ionic compounds</u>, the cation, a metal, is written first and the anion ends in -ide if it is a monatomic ion (ion containing one atom).
  - For example, the halogen ions are fluoride, chloride, bromide, and iodide.



the electron lost by a sodium atom is gained by a chlorine atom to produce a sodium ion and a chloride ion



Oppositely charged ions then attract each other, that attraction called an <u>ionic bond</u>.

Ions Can Be Molecules *(charged molecules)*

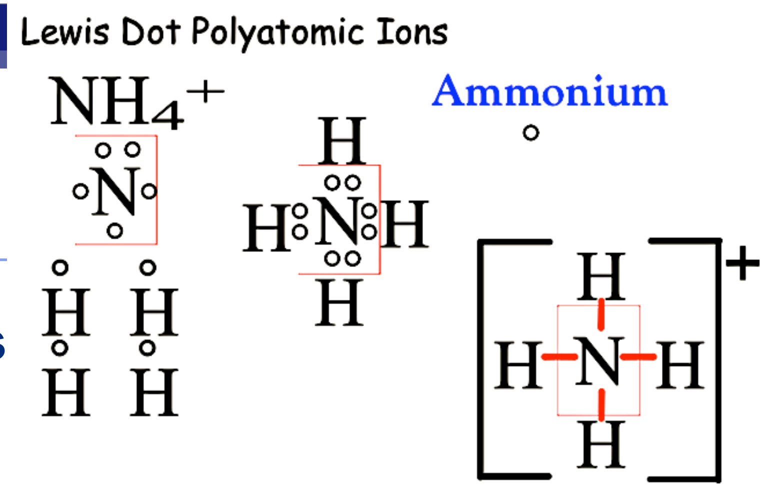
- **Molecules are groups of two or more atoms covalently bonded together that form the smallest identifiable unit into which a pure substance can be divided.**
  - Molecules can be neutral *(all charges in all atoms sum to 0 net charge)* or carry a net charge.
- **Polyatomic ions are covalently bonded molecules that carry a net charge due to an imbalance in the total number of protons versus electrons in the molecule.**
  - Polyatomic ions are similar to monatomic ions, in that they have positive or negative net charge, which is due to the gaining or loss of electrons.
- **A polyatomic ion will attract an ion of opposite charge**
  - Ionic compounds have an overall neutral charge, meaning the ions that come together to form a unit with a net neutral charge.
    - To write the molecular formula for a compound with polyatomic ions...
      - First write the formula and charge for the cation / polyatomic cation
      - Next write the formula and charge of the anion / polyatomic anion.
      - Balance the charges by using parentheses if there is more than one polyatomic ion.
        - For example, $MgSO_4$ versus $Ni(SO_4)_2$.
        - For example, NaOH versus $Ca(OH)_2$.

SO_4^{2-}

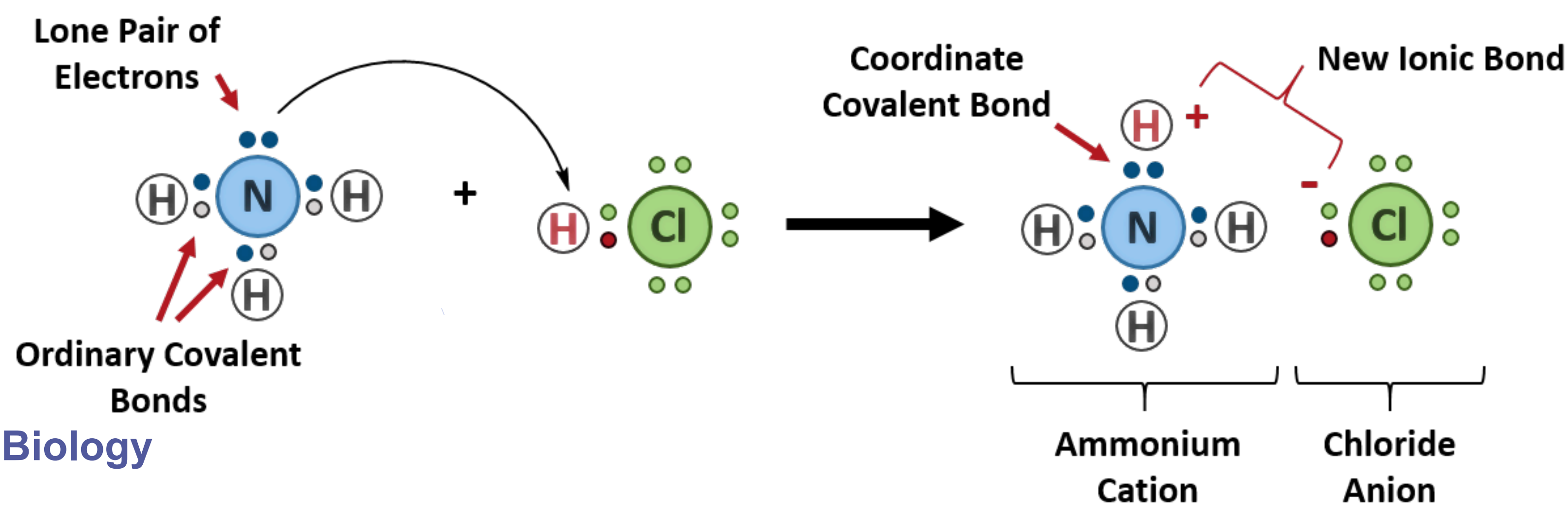
OH^-

Ca^{2+} Mg^{2+}

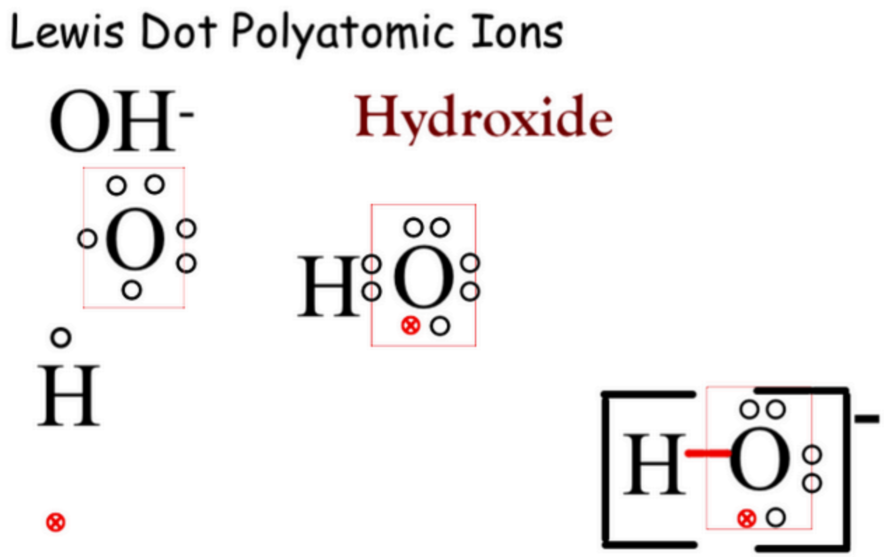
Polyatomic Ions



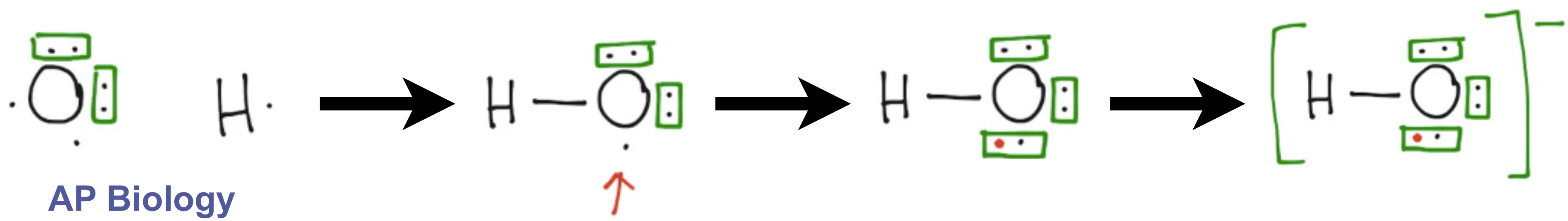
- Nitrogen can bond with three hydrogens covalently to make ammonia (NH_3), an uncharged, yet polar molecule.
- However, because of the lone pair of electron on the nitrogen in ammonia, N can also form a covalent bond with an additional hydrogen that has lost its own electron (an H+ ion or a proton).
  - This creates a molecule with an electrochemical imbalance of +1, the positive polyatomic ammonium, NH_4^+
    - Because the molecule shown below is polyatomic, we place the structure in brackets and place the charge as a superscript outside the bracket.



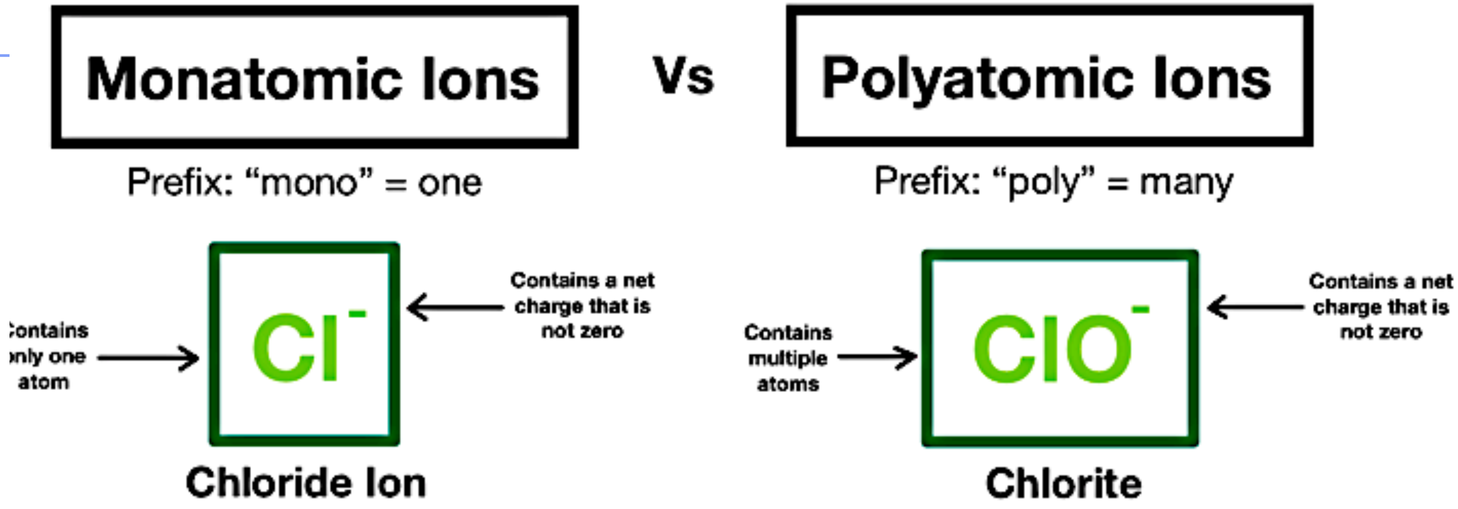
Polyatomic Ions



- For the negative polyatomic ion, hydroxide, the single hydrogen can only covalently bond to one of the bond sites on the oxygen.
- This leaves an empty bond site which can be **filled by an atmospheric electron** to complete the rule of octet for the oxygen.
  - The addition of the additional electron creates a molecule with an electrochemical imbalance of -1, the negative polyatomic ion **hydroxide, OH^-**.
    - Because it is a polyatomic we place the structure in brackets and place the charge as a superscript outside the bracket.



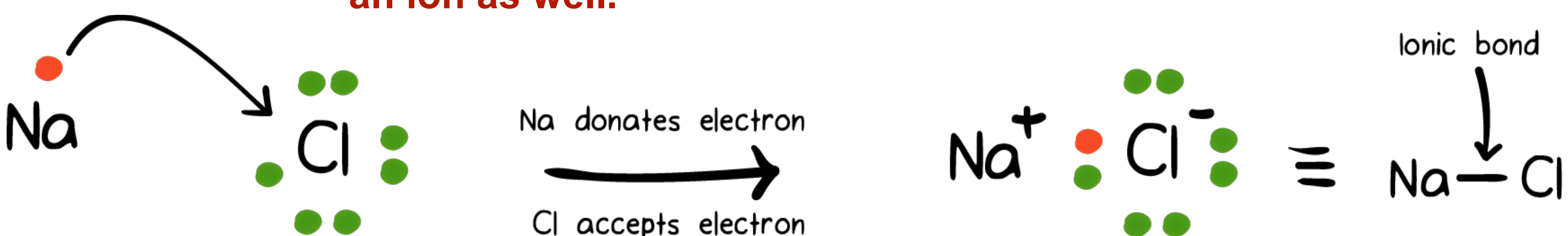
Polyatomic Ions



Ion	Name	Ion	Name
NH_4^+	Ammonium	CO_3^{2-}	Carbonate
NO_2^-	Nitrite	HCO_3^-	Bicarbonate
NO_3^-	Nitrate	ClO_4^-	Perchlorate
SO_3^{2-}	Sulfite	ClO_3^-	Chlorate
SO_4^{2-}	Sulfate	ClO_2^-	Chlorite
HSO_4^-	Bisulfate	ClO^-	Hypochlorite
OH^-	Hydroxide	$C_2H_3O_2^-$	Acetate
PO_4^{3-}	Phosphate	$Cr_2O_7^{2-}$	Dichromate
HPO_4^{2-}	Hydrogen Phosphate	CrO_4^-	Chromate
$H_2PO_4^-$	Dihydrogen phosphate	CN^-	Cyanide

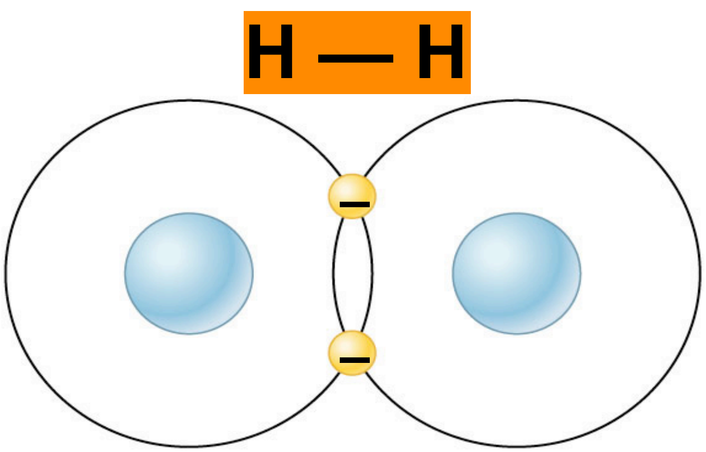
Type of Intramolecular Forces of Attraction = Ionic Bonds

- **Ionic bond: This strong bond is formed after the complete transfer of valence electron(s) between atoms leads the atoms (or molecules) to form oppositely charged ions (or polyatomic ions) that then attract one another.**
  - **To form an ionic bond, the metal atom loses electrons to become a positively charged cation, whereas the nonmetal atom accepts those electrons to become a negatively charged anion, the opposite charges then attracting.**
    - **The difference in electronegativity between bonded atoms is over 2.0.**
      - **Ionic bonds can form between polyatomic ions and an ion as well.**

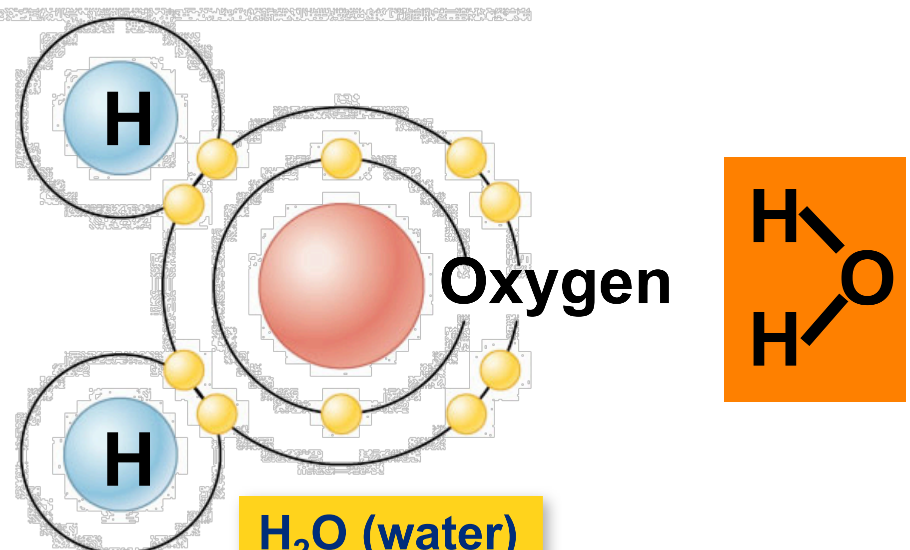


Type of Intramolecular Forces of Attraction = Covalent Bonds

- **Ionic bonds** form between a metal atom and a nonmetal atom (atoms very different in Electronegativity).
  - But what happens when the two atoms are closer in electronegativity as is the case when two nonmetals meet?
- Why are **covalent bonds** strong bonds?
  - two atoms “share” a pair of electrons
  - both atoms’ positively charged nuclei pull & hold onto the bonding electrons
  - Since they are strong bonds, they are very stable
- **Molecules** are groups of atoms joined via covalent bonds



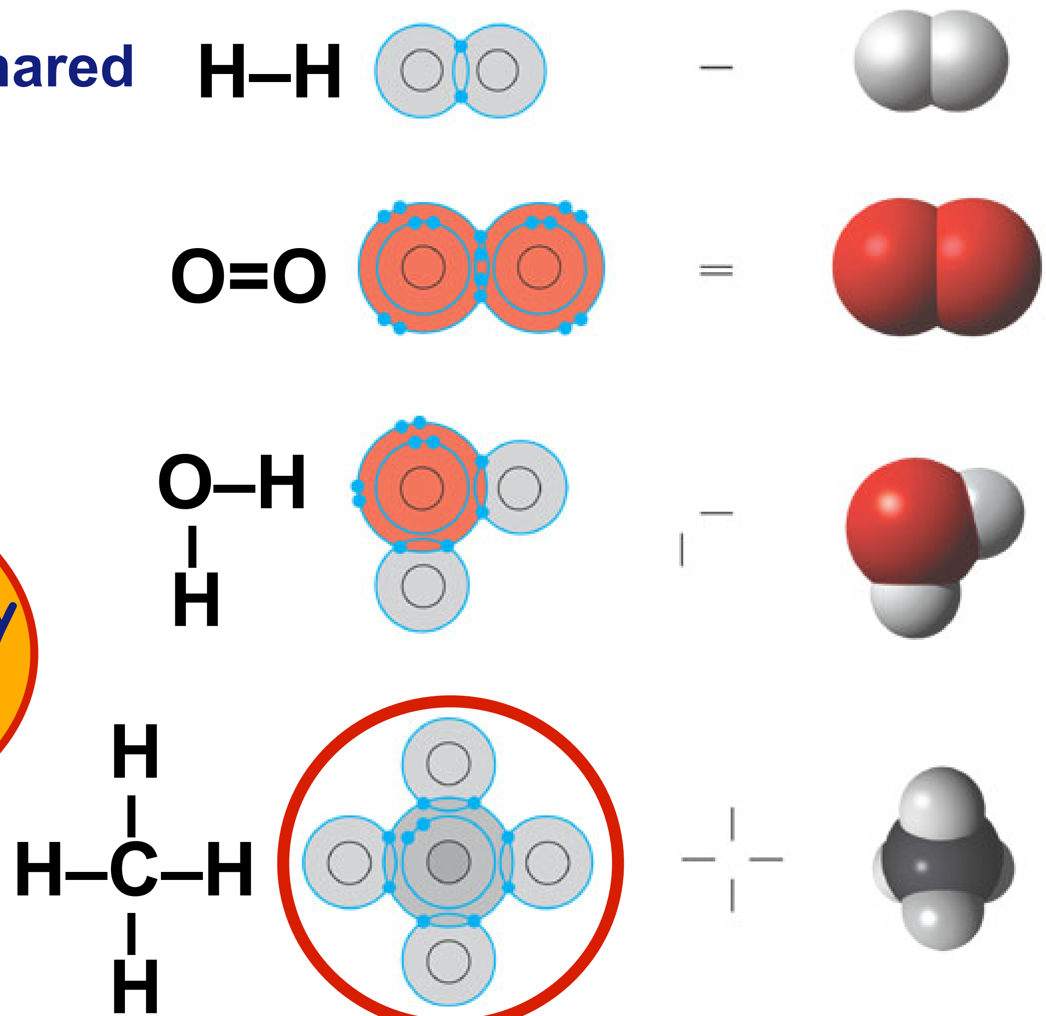
H_2 (hydrogen gas)



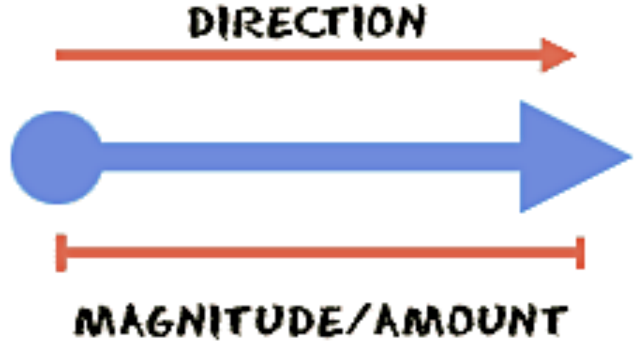
H_2O (water)

Multiple Covalent Bonds Can Exist between Atoms

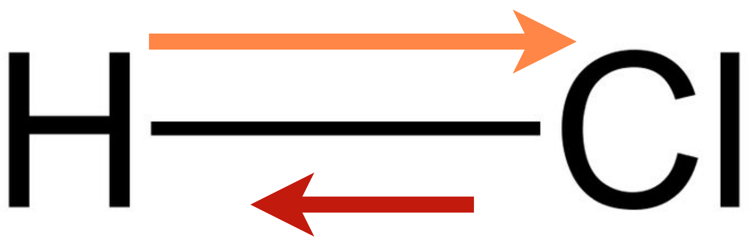
- 2 atoms can share more than 1 pair of electrons
  - single bonds
    - 1 pair of electrons shared between 2 atoms
  - double bonds
    - 2 pairs of electrons
  - triple bonds
    - 3 pairs of electrons



Vectors Represent Forces

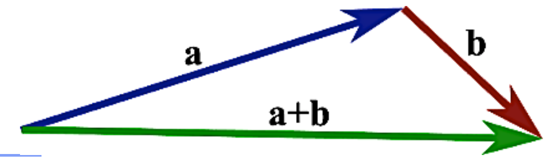


- Covalent bonds involve the nuclei of two atoms exerting forces of attraction on the valence electrons being "shared."
  - These forces have defined directions (in 3D space) & magnitudes that can be symbolized by arrows called vectors.
    - The vector's arrow head points in the direction of the force of attraction, towards the nucleus, which is electronegative.
    - The length of the vector indicates the strength of the force, the pull by the nucleus on electrons.

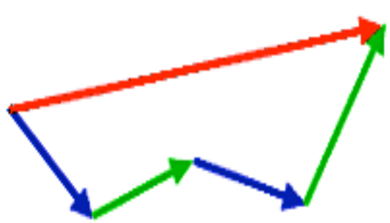


      - In the example of the covalent bond between H and Cl in HCl, the protons of the Cl atom exert a force of attraction on the electron of the H atom.
      - At the same time, the protons of the H atom exert a force of attraction on the electron of the Cl atom.

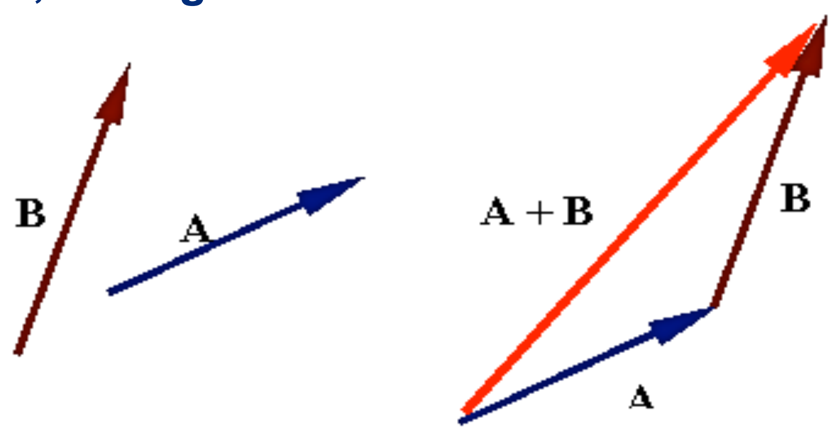
Vectors can be Added

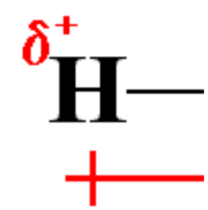


- **Forces within a bond or within a molecule with many covalent bonds can be summed.**
  - **To add vectors, we chain the vectors by putting the tail of the next vector on the head of the previous vector.**
    - **The overall net vector is then drawn from the tail of the first vector to the head of the last vector**
      - **This resultant net vector is the <u>permanent dipole moment</u> in the bond or in the molecule.**
        - **When summing forces, they may cancel each other out, resulting in <u>NO NET</u> (overall) force.**
        - **They may not cancel each other out resulting in a <u>NET</u> force with a certain Direction & Magnitude.**
          - **Ex: Below, adding vectors A and B results in the vector A+B**



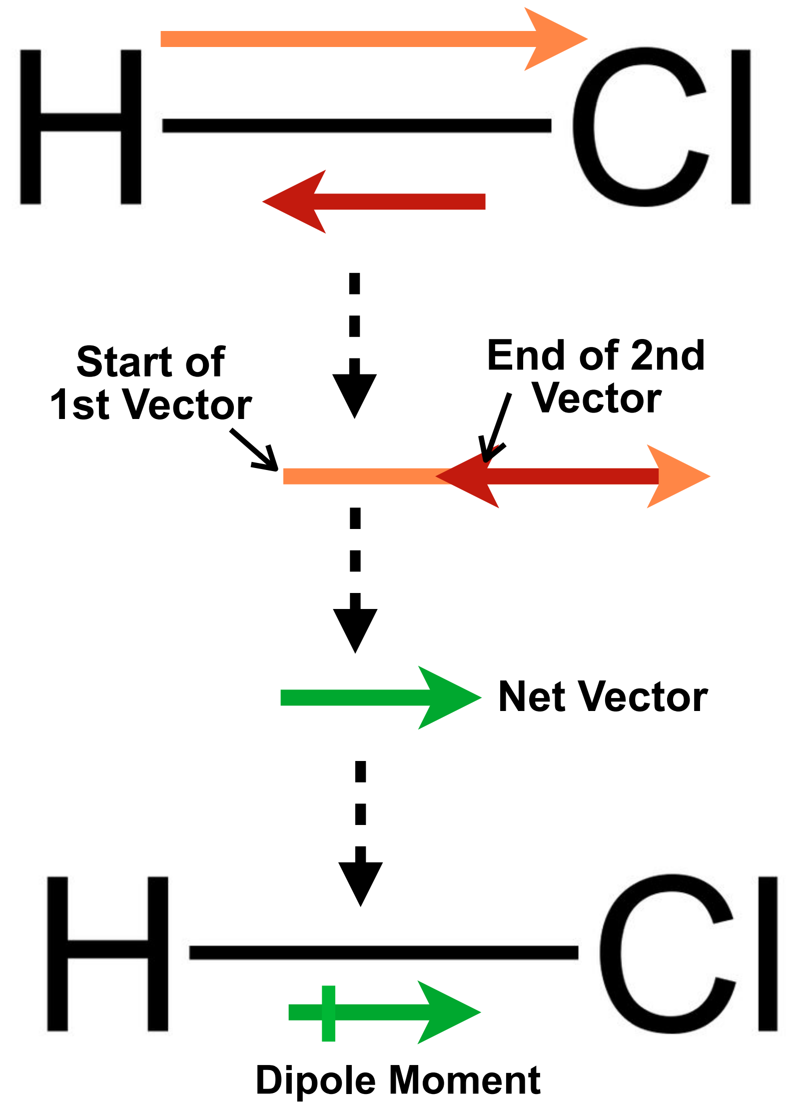
THE RED VECTORS ARE THE RESULT OF ADDING THE SMALLER COLORED ONES.

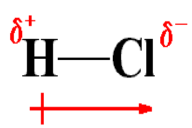




Adding the Forces of Attraction Between Two Atoms

- To add the forces being exerted on the bonding electrons between H and Cl, place the vectors head to tail.
  - Since the head of the last vector does not end up exactly where the tail of the first vector originated, the forces do <u>not</u> fully cancel each other out.
    - A <u>net bond vector (a dipole moment)</u> exists in the H to Cl covalent bond.
      - ➡ The tail of the net vector starts at the origin of the tail of the 1st vector.
      - ➡ The head of the net vector ends right where the head of the last vector added is located.
        - Electron density is pulled with greater force towards the Cl than towards the H.

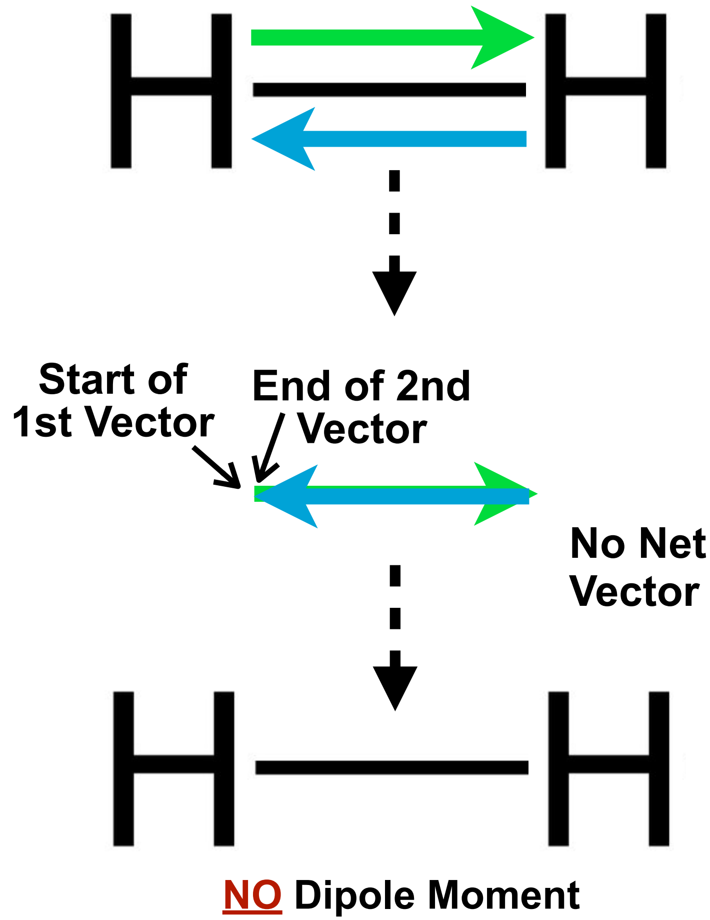


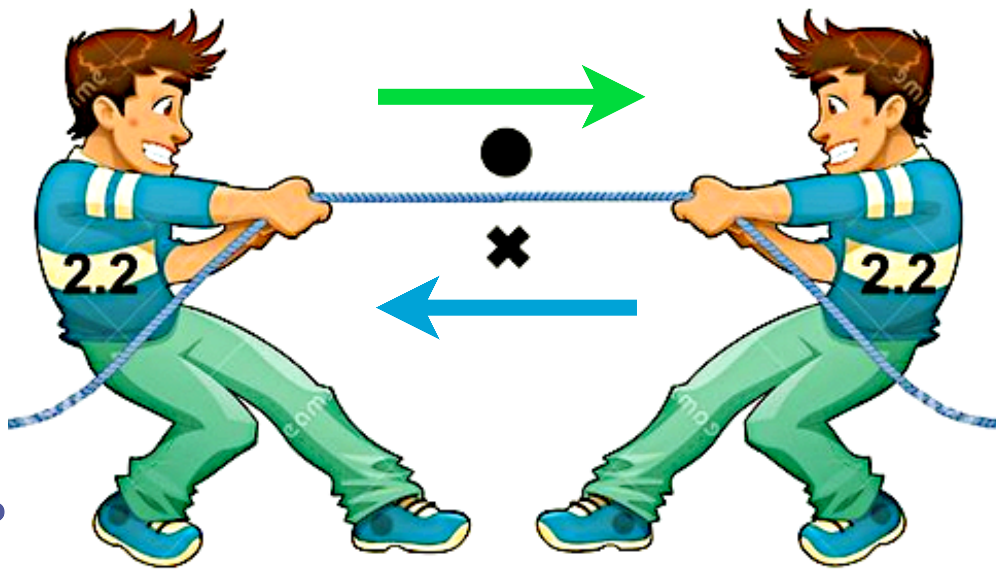


Dipole Moment has a Magnitude and a Direction

Adding the Forces of Attraction Between Two Atoms

- In the case of H_2, both atoms are **equally electronegative**.
- The nuclei of both H atoms exert an equal and opposite force of attraction on the electrons of the neighboring H atom.
  - Since the head of the last vector does end exactly where the tail of the first vector originated, the forces DO fully cancel each other out.



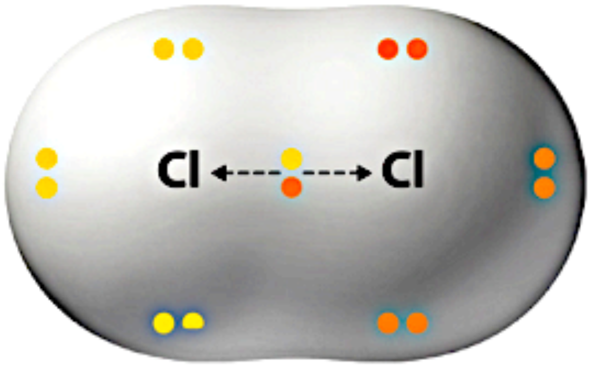


- Electron density is NOT pulled with greater force towards one H or the other H.

AP

Polar vs Non-Polar Covalent Bonds

- When forces in a covalent bond are added, they can either cancel each other out (result in NO net vector) or result in a NET vector, referred to in chemistry as a bond dipole moment.
  - If the forces do cancel each other out, the electrons being shared, are shared equally.
    - The bond is considered non-polar.
    - The vectors cancel each other out, leaving no dipole moment (net force).
    - No dipole forms (there are *no* equal and oppositely charged poles separated by a distance across this bond)



Nonpolar covalent bond

Bonding electrons shared equally between two atoms. No charges on atoms.

Non-Polar Covalent Bonds

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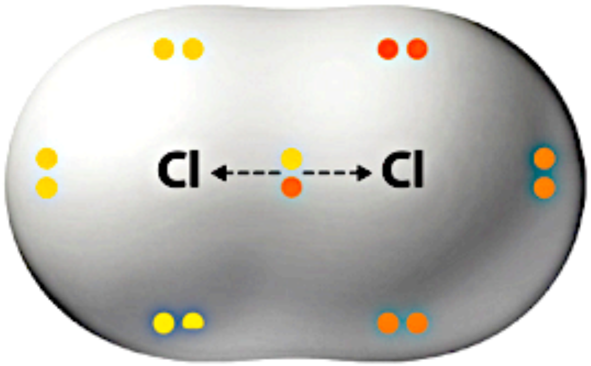
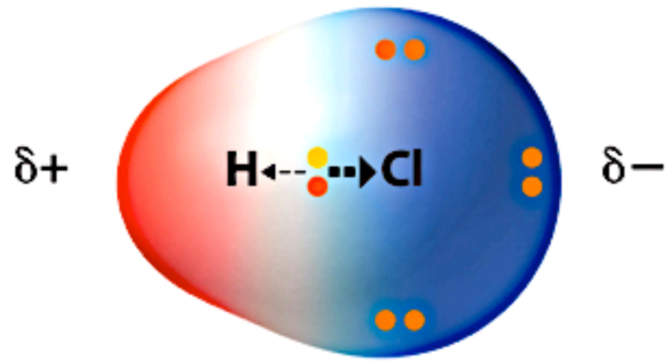


Nonpolar covalent bond

Bonding electrons shared equally between two atoms. No charges on atoms.

Polar vs Non-Polar Covalent Bonds

- **When forces in a covalent bond are added, they can either cancel each other out (result in NO net vector) or result in a NET vector, referred to in chemistry as a bond dipole moment.**
- **If the forces do not cancel each other out, the electrons being shared, are shared unequally.**
  - **The bond is considered polar.**
  - **The vectors do not cancel each other out, leaving an overall dipole moment**
  - **A dipole forms (there *are* equal and oppositely charged poles separated by a distance across this bond)**
    - **We indicate that the bonding electrons spend more time in the electron cloud of Cl than that of H (example shown) by placing partial - and partial + charge symbols on Cl and H, respectively..**

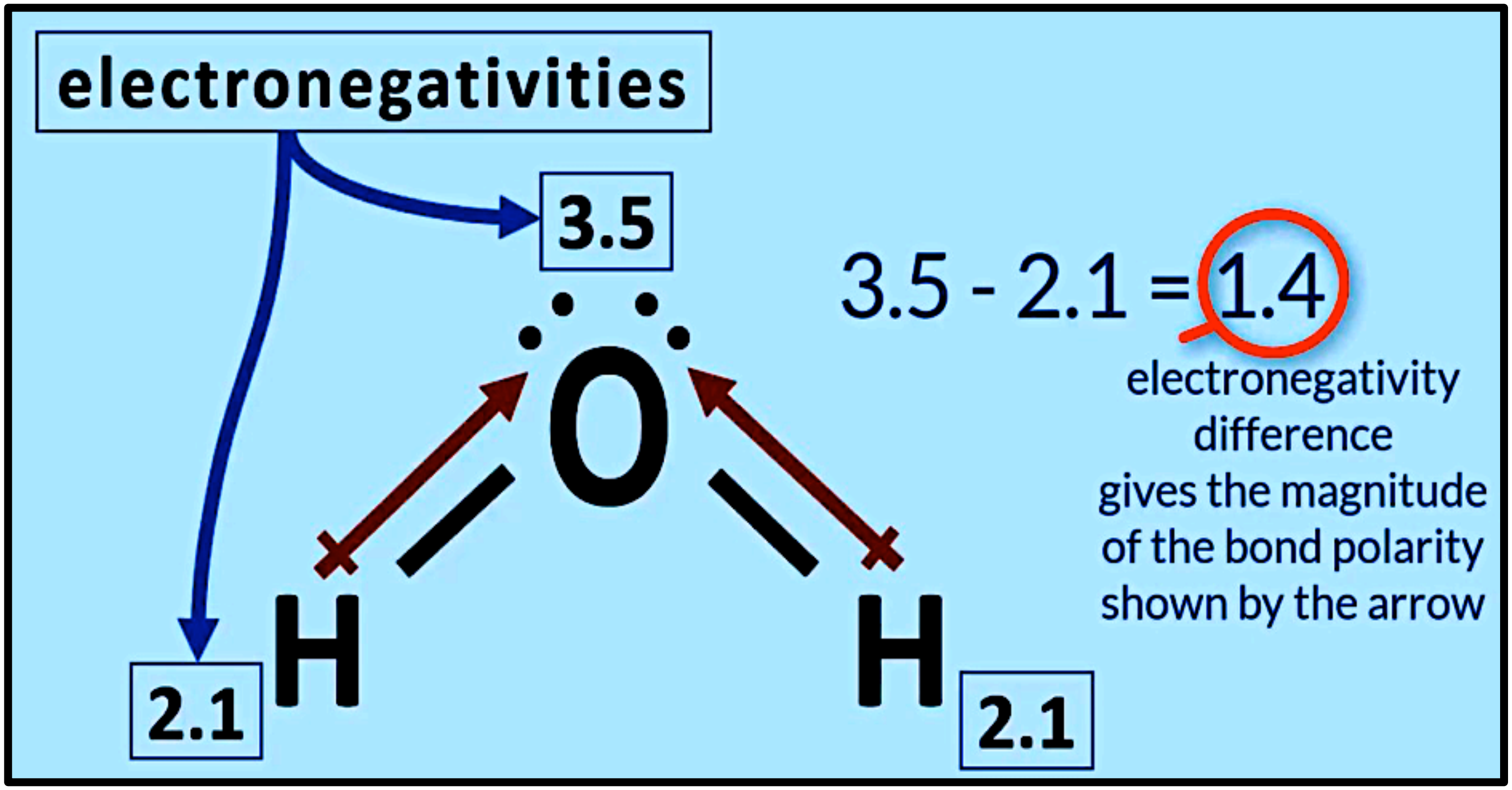


Polar covalent bond

Bonding electrons shared unequally between two atoms. Partial charges on atoms.

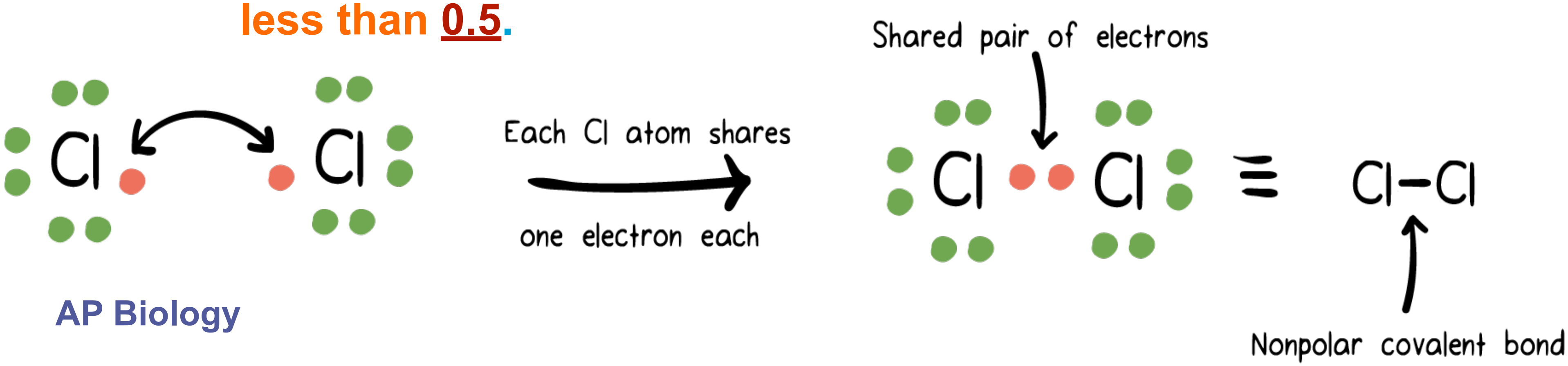
Polar vs Non-Polar Covalent Bonds

- **Whether a bond is non-polar or polar thus also depends on the DIFFERENCE in Electronegativity between the two atoms involved in the covalent bond.**

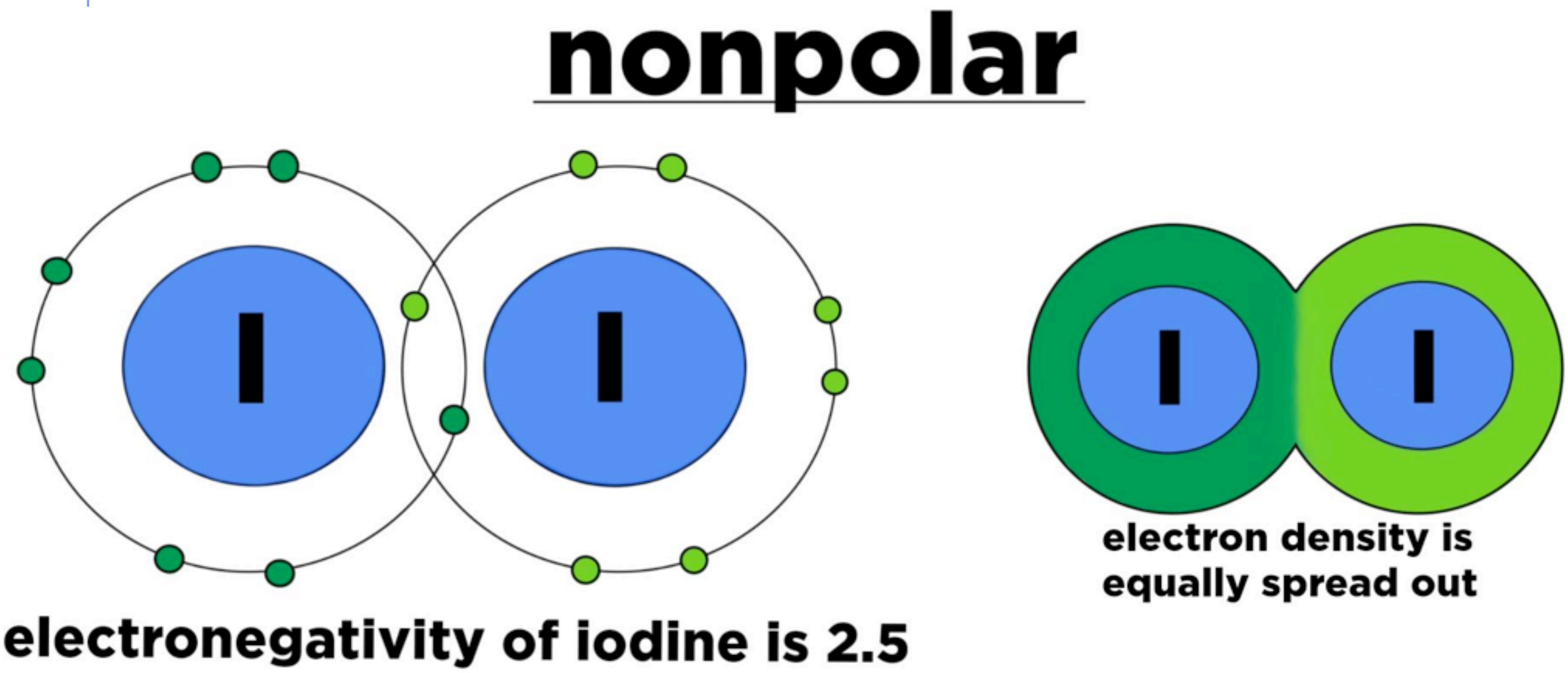


Type of Intramolecular Forces of Attraction = Nonpolar Covalent Bonds

- Covalent bonds forms between atoms that have more similar electronegativities
- Because both atoms have a more similar affinity for electrons, neither has a tendency to lose or permanently gain them.
  - The electrons are shared in order to achieve octet configuration whenever possible and so to become more stable.
- A nonpolar covalent bond is formed between same atoms or atoms with very similar electronegativities.
  - The difference in electronegativity between bonded atoms is less than 0.5.

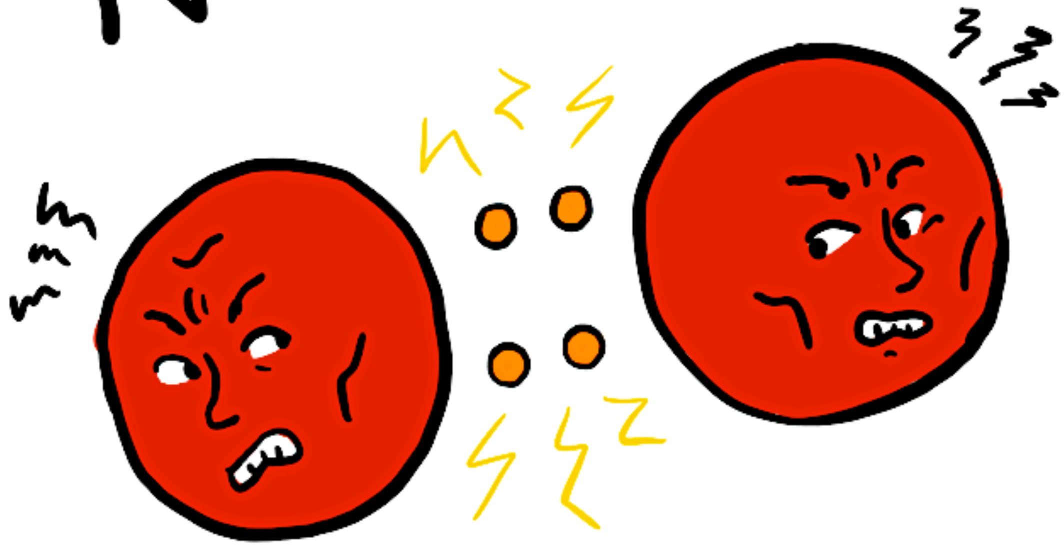


Non-Polar Covalent Bonds



Non-Polar Covalent Bonds

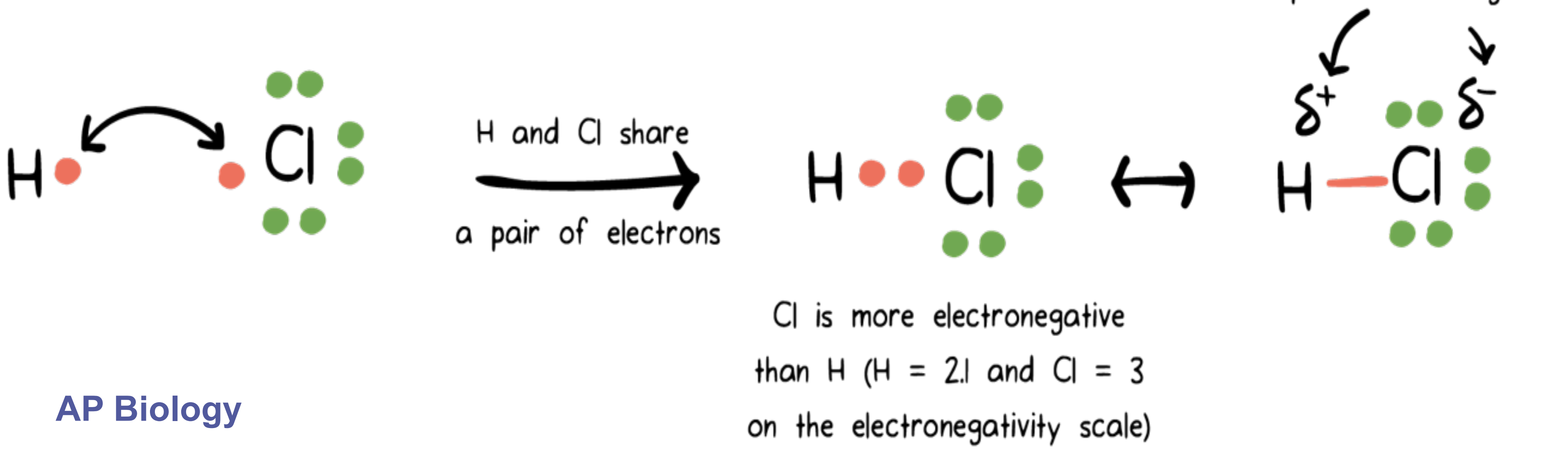
NonPolar: O_2



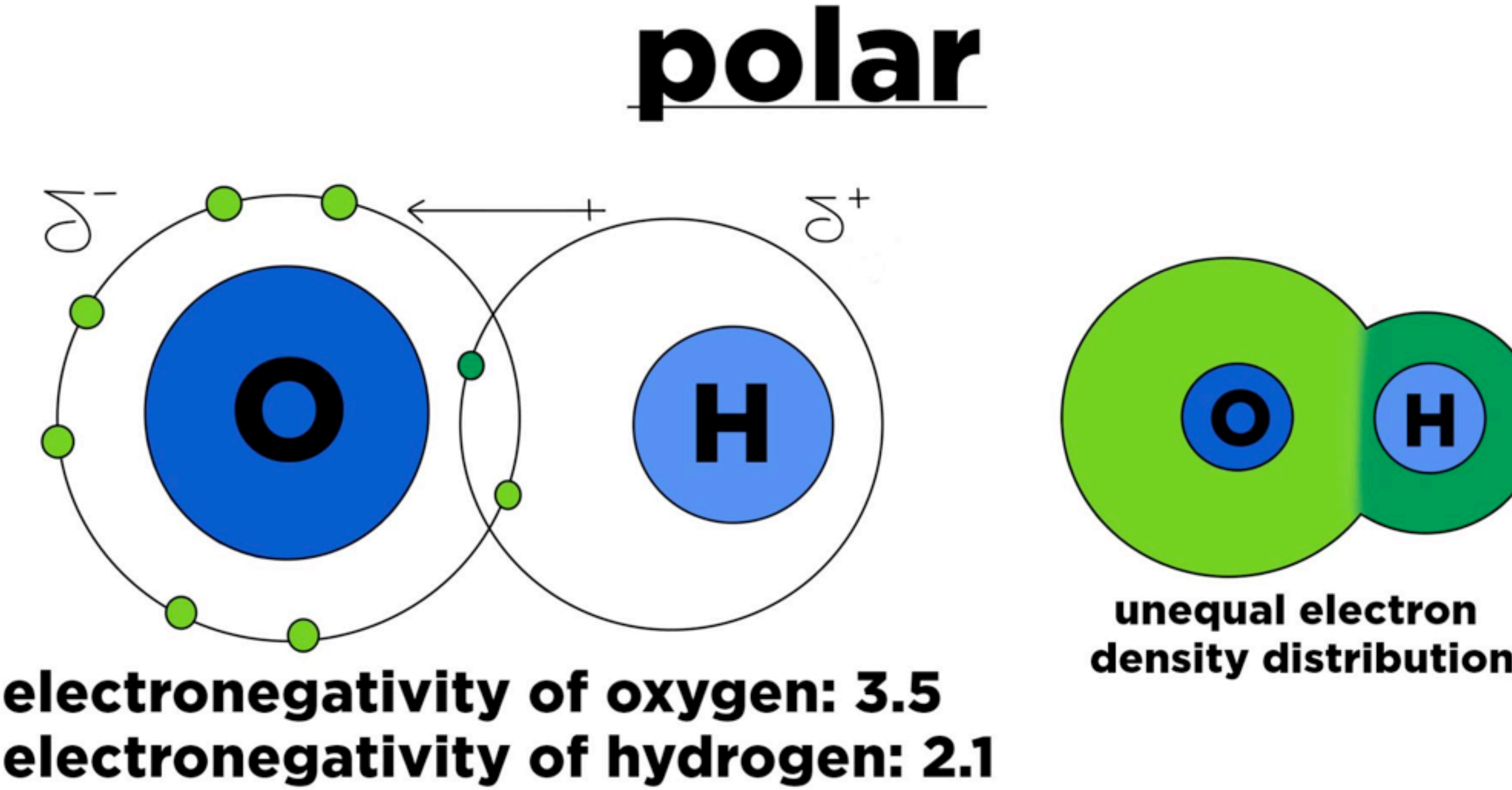
Although oxygen is very electronegative, O_2 (O=O) is **not** Polar. Because both atoms have the same electronegativity, electrons are shared equally between them.

Type of Intramolecular Forces of Attraction = Polar Covalent Bonds

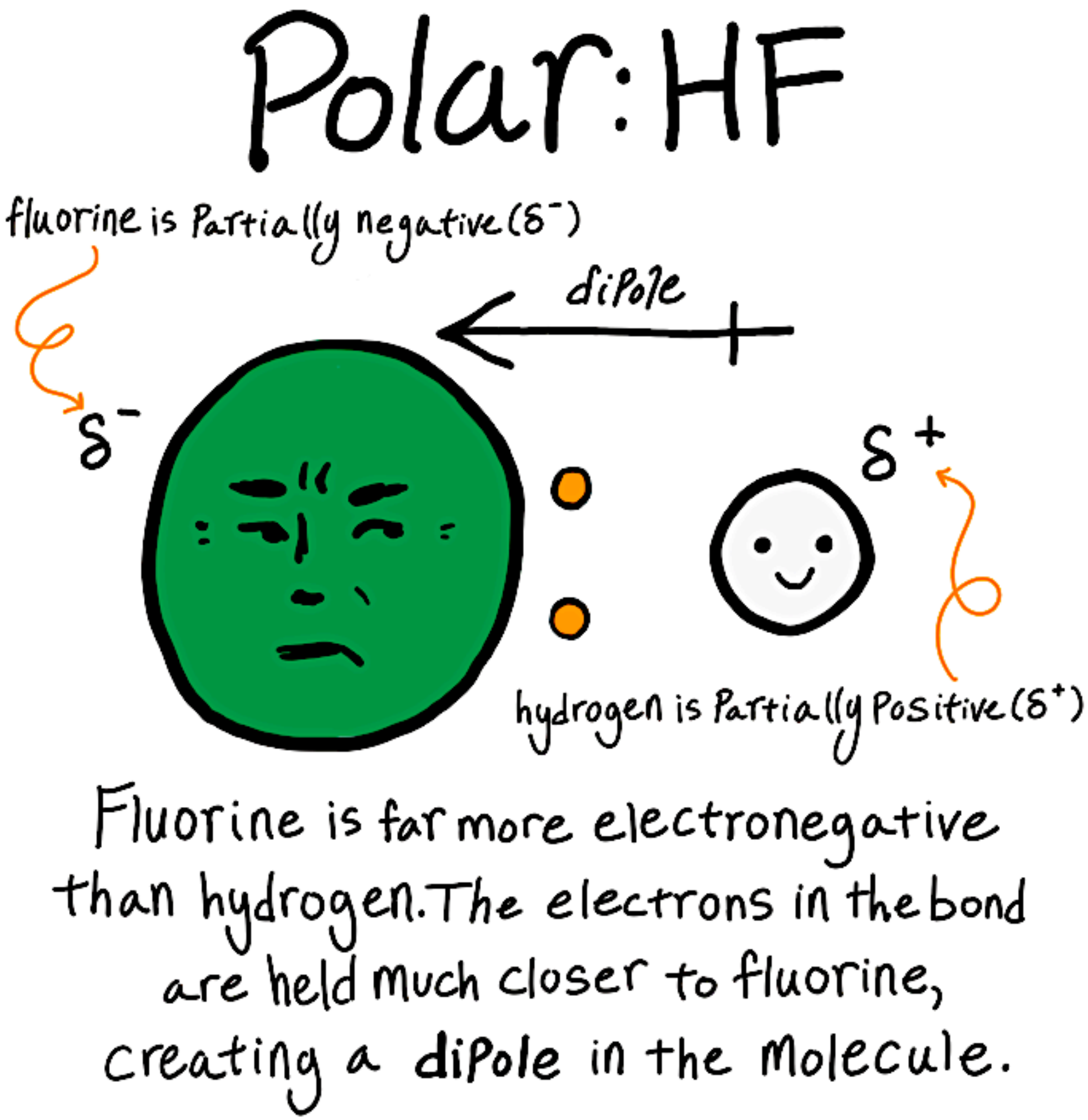
- A polar covalent bond is formed between atoms with slightly larger differences in electronegativities.
  - The difference in electronegativity between bonded atoms is between 0.5-1.9.



Polar Covalent Bonds

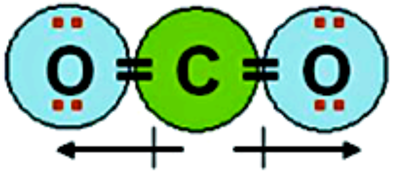


Polar Covalent Bonds



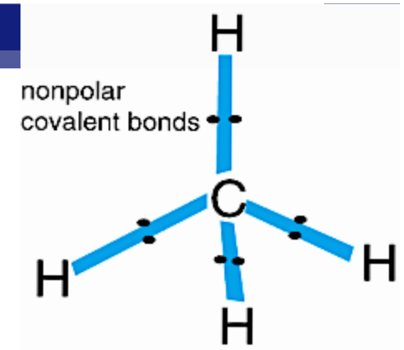
Molecules too can be nonpolar or polar not just individual covalent bonds!

- Just like covalent bonds can be polar or nonpolar, molecules (with one or more covalent bonds) can be nonpolar or can be polar
  - Nonpolar molecules have no dipoles
    - A molecule with more than one polar bond might not have a permanent dipole moment when the charge separations are symmetrically distributed so the net vectors of all covalent bonds sum up to 0.
  - Polar molecules have one or more dipoles and overall net molecular dipole moment.
    - A molecule with more than one polar bond might have a permanent dipole moment when the net vectors of all covalent bonds do not cancel each other.

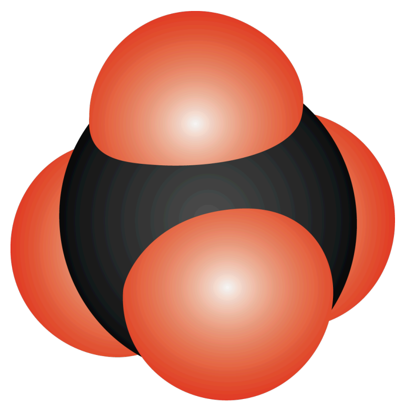


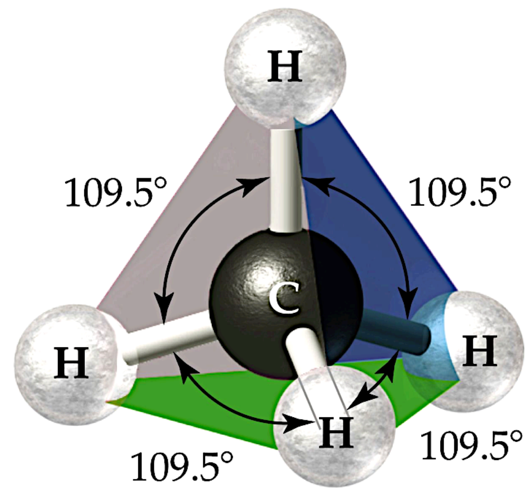
dipole moment = 0

Molecules with No Polar Bonds



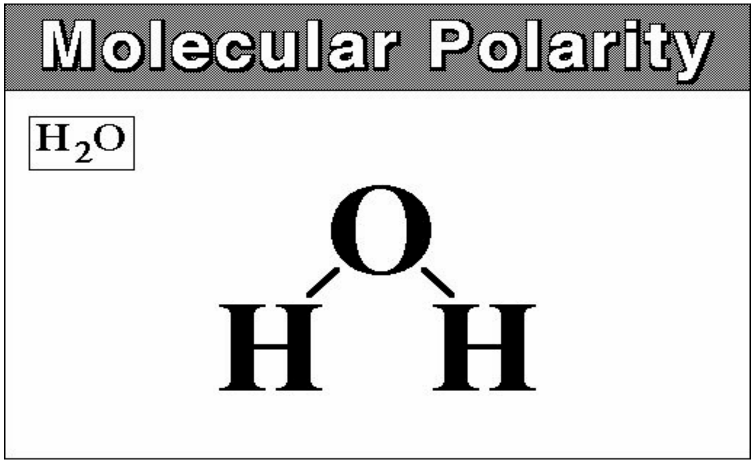
- Electrons in each covalent bond are **<u>shared fairly equally</u>** by 2 atoms - the atoms do <u>not</u> acquire significant opposite net charges (there are <u>no</u> dipoles)
  - *If all the covalent bonds in a molecule are non-polar, the molecule as a whole is <u>non-polar</u> (not a dipole)*
    - <u>Example:</u> Generic hydrocarbons = C_xH_x
      - Specific Ex: methane (CH_4)



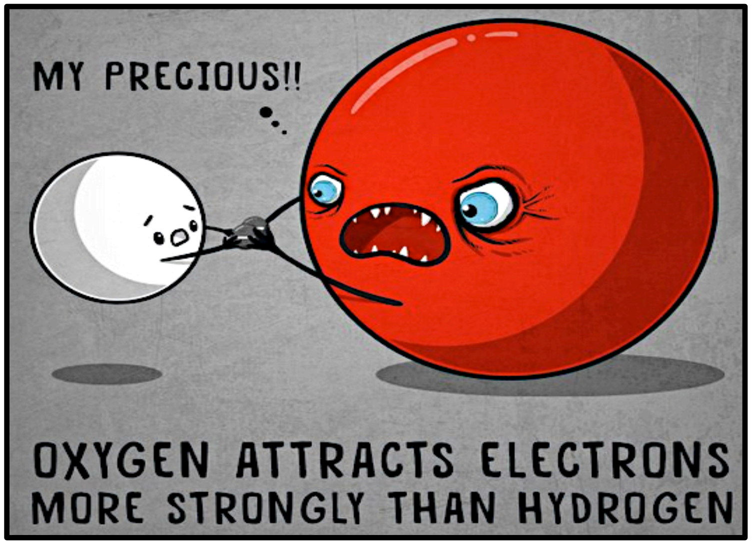


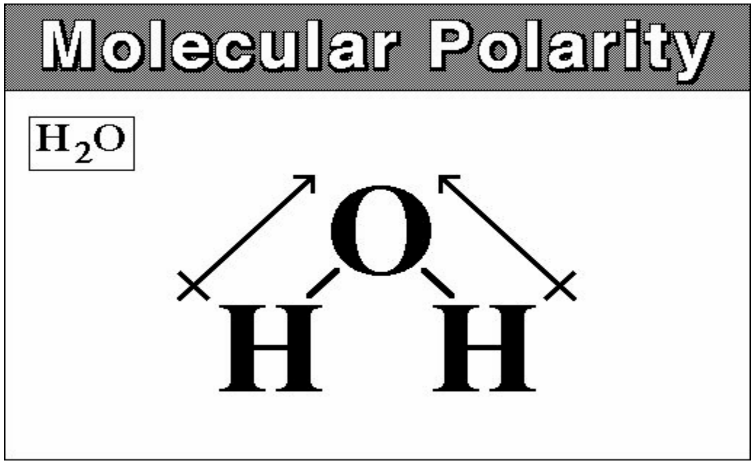
Nonpolar molecule that doesn't regularly attract molecules of itself easily. Has no "permanent" partial charges on any atom.

Why H_2O is a POLAR MOLECULE with POLAR BONDS?

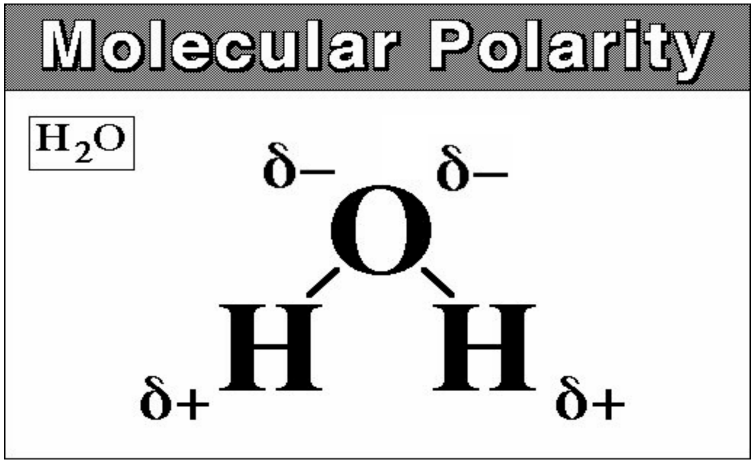


1. H_2O is a **bent** molecule



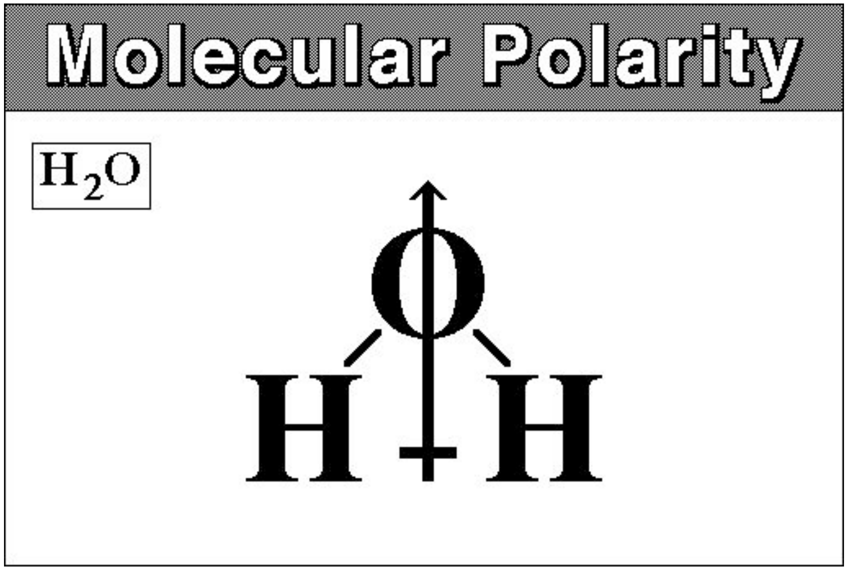


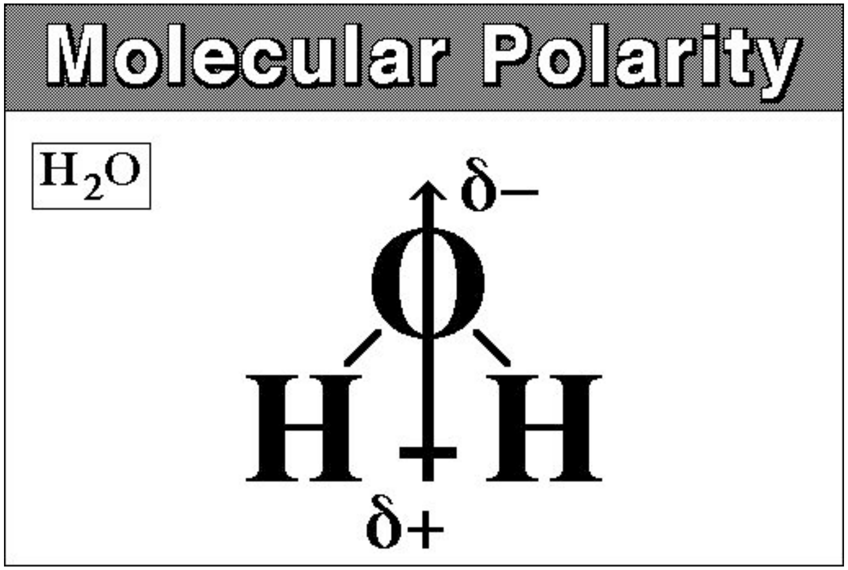
2. O = high E.N. & H = "medium" E.N. so the two covalent bonds are polar with a (net force) **dipole moment** associated with each bond



3. The two dipole moments lead to **four partial charged regions** on the molecule.

Why H_2O is a POLAR MOLECULE with POLAR BONDS?



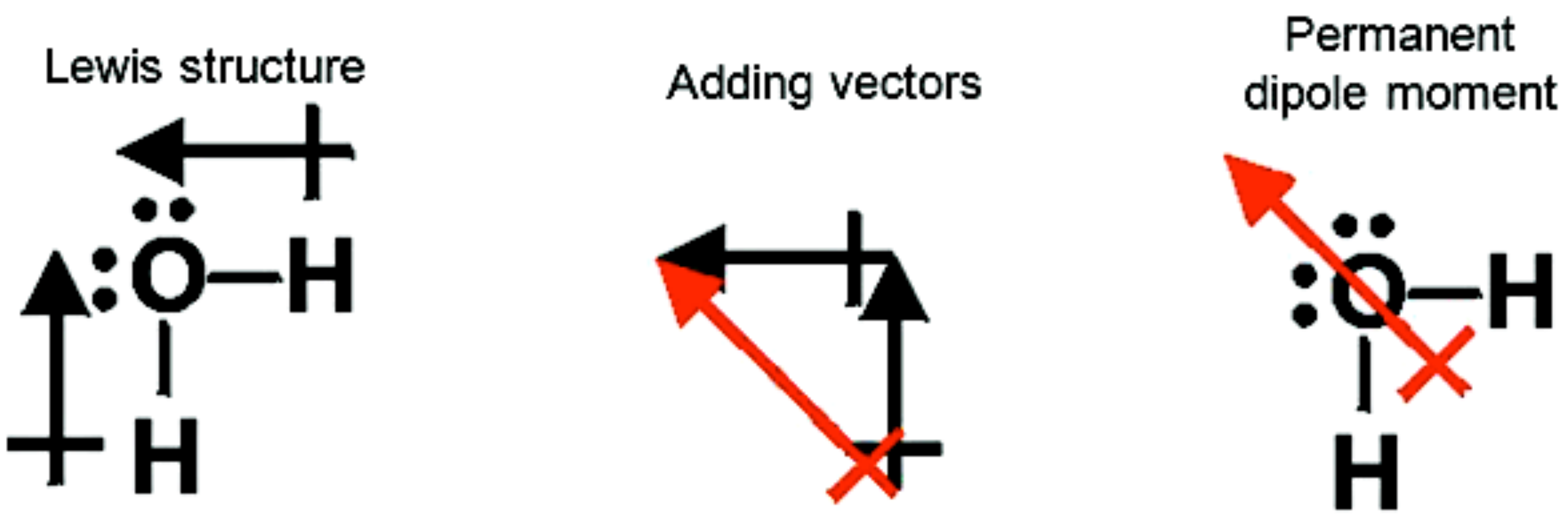


4. Summing the individual dipole moments from ***all*** polar bonds, results in an <u>overall net MOLECULAR dipole moment</u>. (The net forces in each bond throughout the molecule do not cancel each other out).

5. Not only does water have one or more polar covalent bonds, but the <u>entire molecule is polar!</u> (Electron density is pulled more strongly and "continually" to one region of the molecule, making one region slightly + and one region slightly - ***most of the time***)

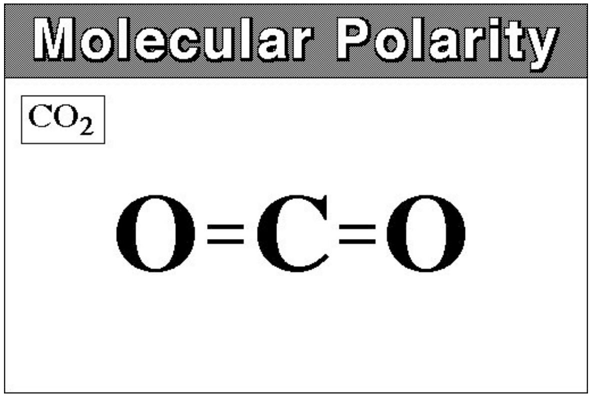
Add **All** Bond Dipole Moments to Determine if the Molecule is Polar

- In water, the two dipole moments (one from each polar covalent bond) do not cancel.
  - The dipole moments (the forces acting on the bonding electrons) result in an overall net MOLECULAR dipole moment.
    - Water molecules are thus a POLAR molecule

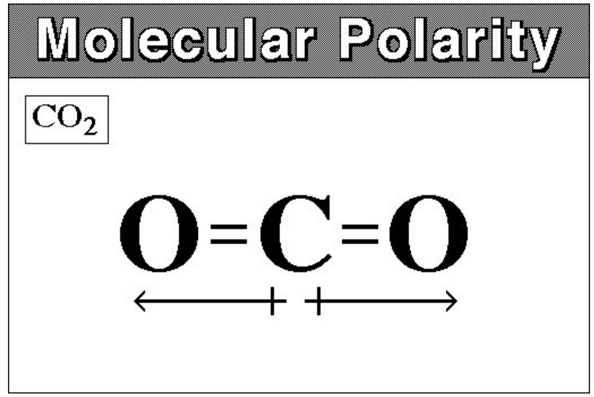


<u>Warning</u>: A molecule is <u>not</u> automatically polar just because it has some polar bonds!

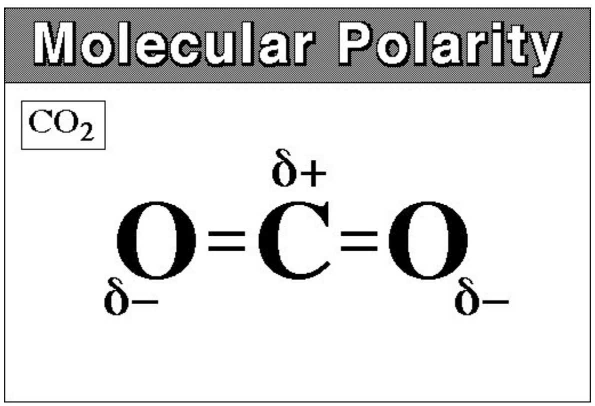
Is CO_2 a <u>POLAR MOLECULE</u> with its <u>POLAR BONDS</u>?



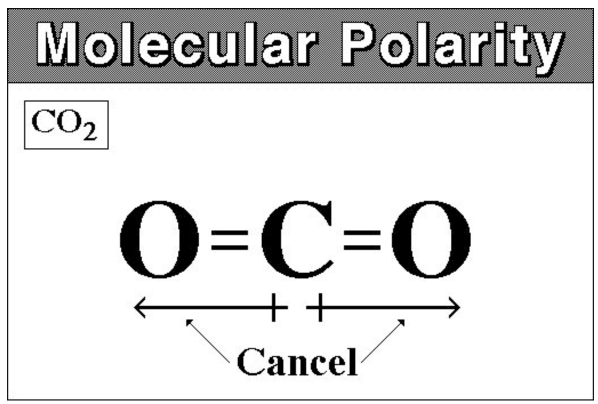
1. CO_2 is a <u>linear</u> molecule



2. O = high E.N. & C = “medium” E.N. so the two two covalent bonds are polar with a (net force) <u>dipole moment</u> associated with each bond



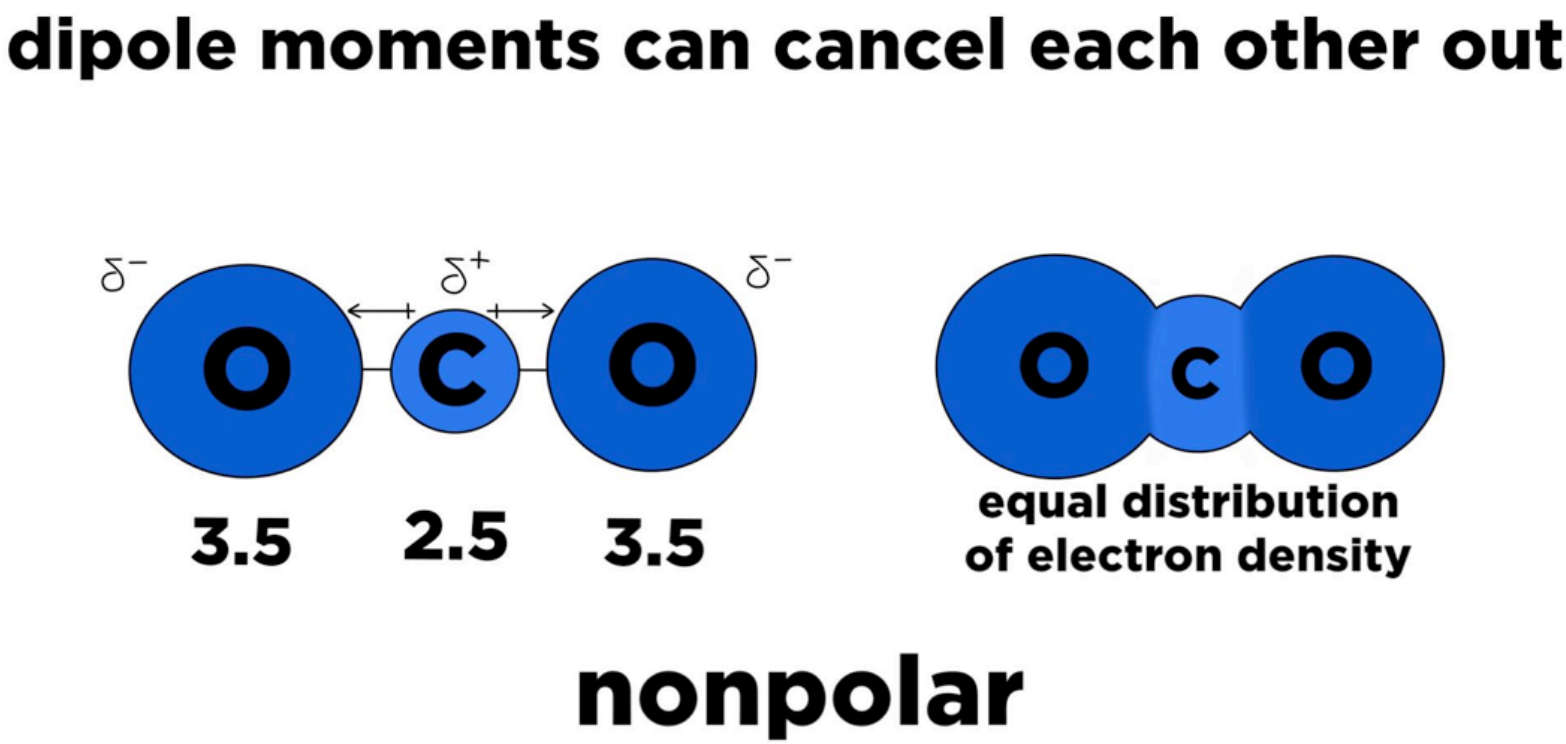
3. The dipole moments lead to <u>partial charged regions</u>



CO_2 is a NONPOLAR molecule

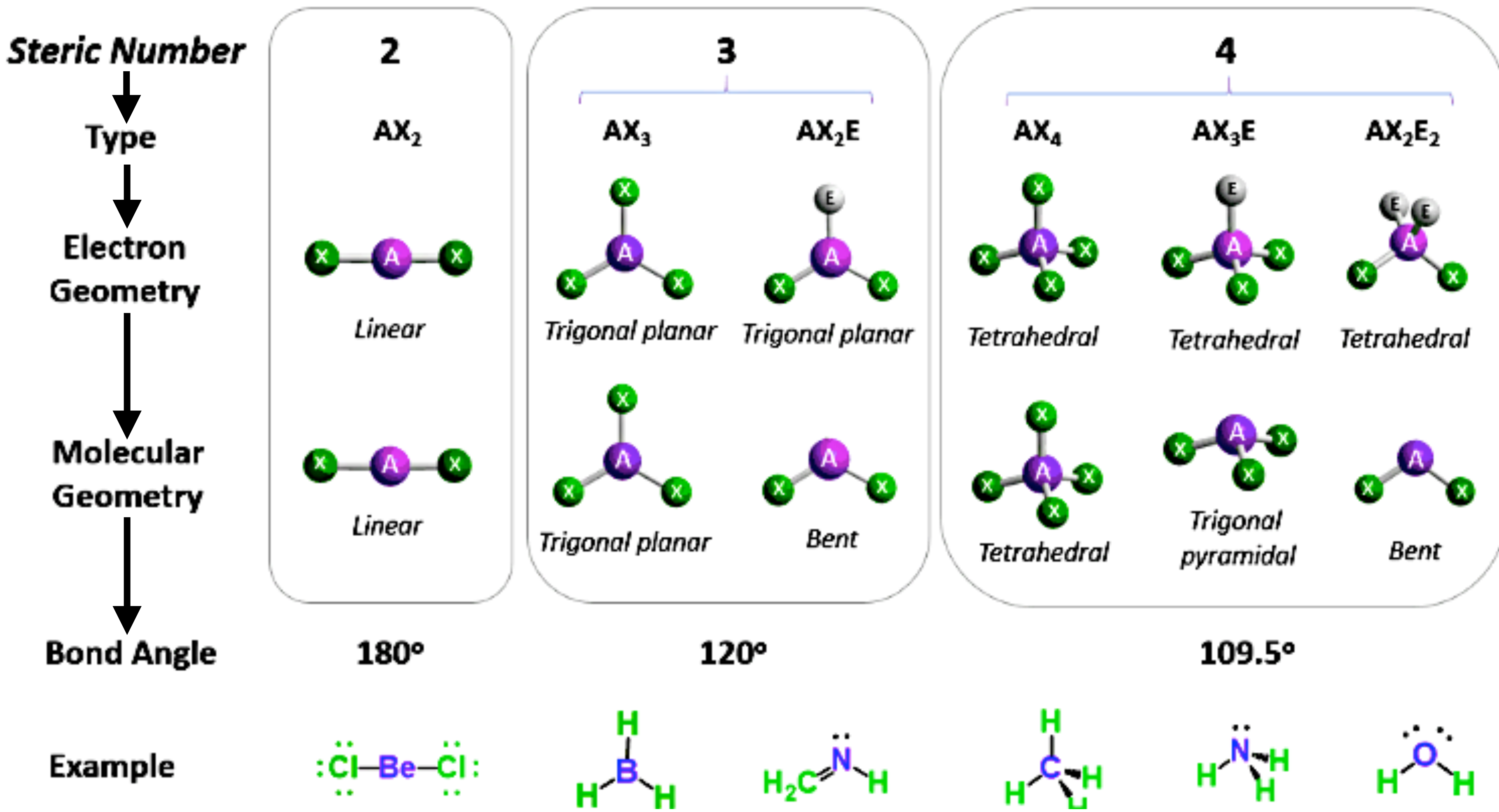
4. Dipole moments cancel when summed, leading to <u>NO</u> <u>MOLECULAR dipole moment</u>.

Is CO_2 a POLAR MOLECULE with its POLAR BONDS?



To determine if dipole moments cancel, one must know the 3D SHAPE of a MOLECULE

The steric number is the number of atoms bonded to a central atom of a molecule plus the number of lone pairs attached to the central atom.



A = Central atom *(most E.N. atom, except Hydrogen which is never A)*

X = An atom or group of atoms

E = Lone pair of nonbonding electrons

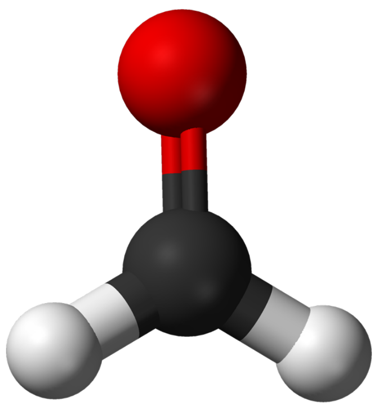
Are there Polar Bonds? Is the Molecule Polar?

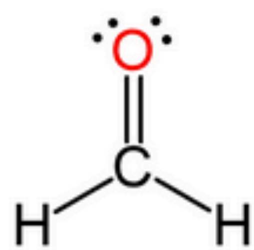
Example:

CH_2O

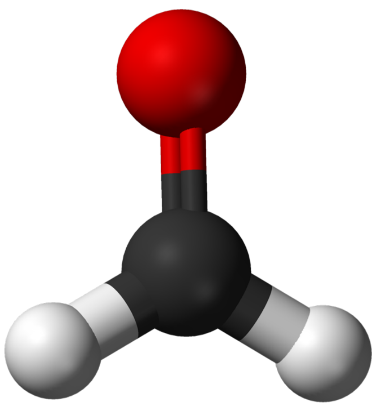
Determine if the molecule is polar or not by showing the corresponding dipole moment(s).

This molecule is formaldehyde and has a **trigonal planar** geometry according to the VSEPR rules:





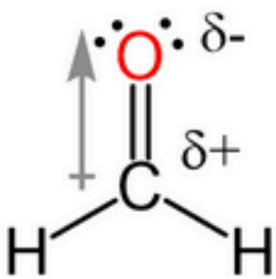
Trigonal planar



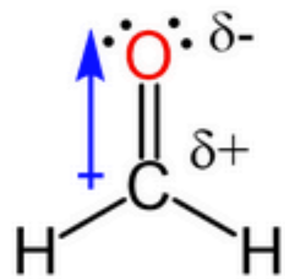
All atoms of this molecule lay within one flat plane with the angles of the three bonds 120 degrees apart from each other

Are there Polar Bonds? Is the Molecule Polar?

The C-H bonds are nonpolar and therefore, the **molecular dipole** is mostly defined by the magnitude and direction of the **C=O bond**. Oxygen, being more electronegative, pulls the electron density of the C=O bond, and thus defining the direction of the molecular dipole:



The dipole moment of the molecule is defined by the C=O bond dipole.



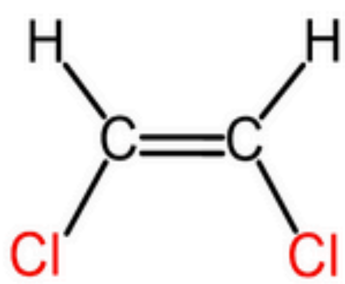
Formaldehyde is a polar molecule since the one dipole moment in the polar C=O bond does not get cancelled out by any other dipole moments from the other bonds, *the other two bonds are nonpolar and so do not have a dipole moment.*

Are there Polar Bonds? Is the Molecule Polar?

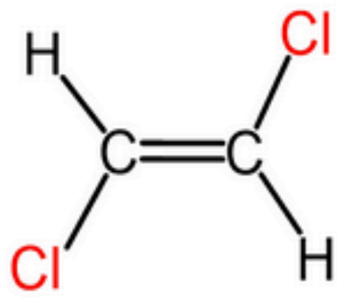
For larger molecules with **lots of sigma bonds**, it is more difficult to determine the net dipole moment because of the **free rotation** about these single bonds. However, the presence of **double bonds** that **restrict** the constant change of the dipole moment direction makes it possible to determine the molecular dipole moment.

This can be illustrated by comparing the boiling points of the ***cis*–** and ***trans*-1,2-Dichloroethene**:

All atoms of these molecule lay within one flat plane.
Where do dipole moments form? Do they cancel each other out?



cis-1,2-Dichloroethene

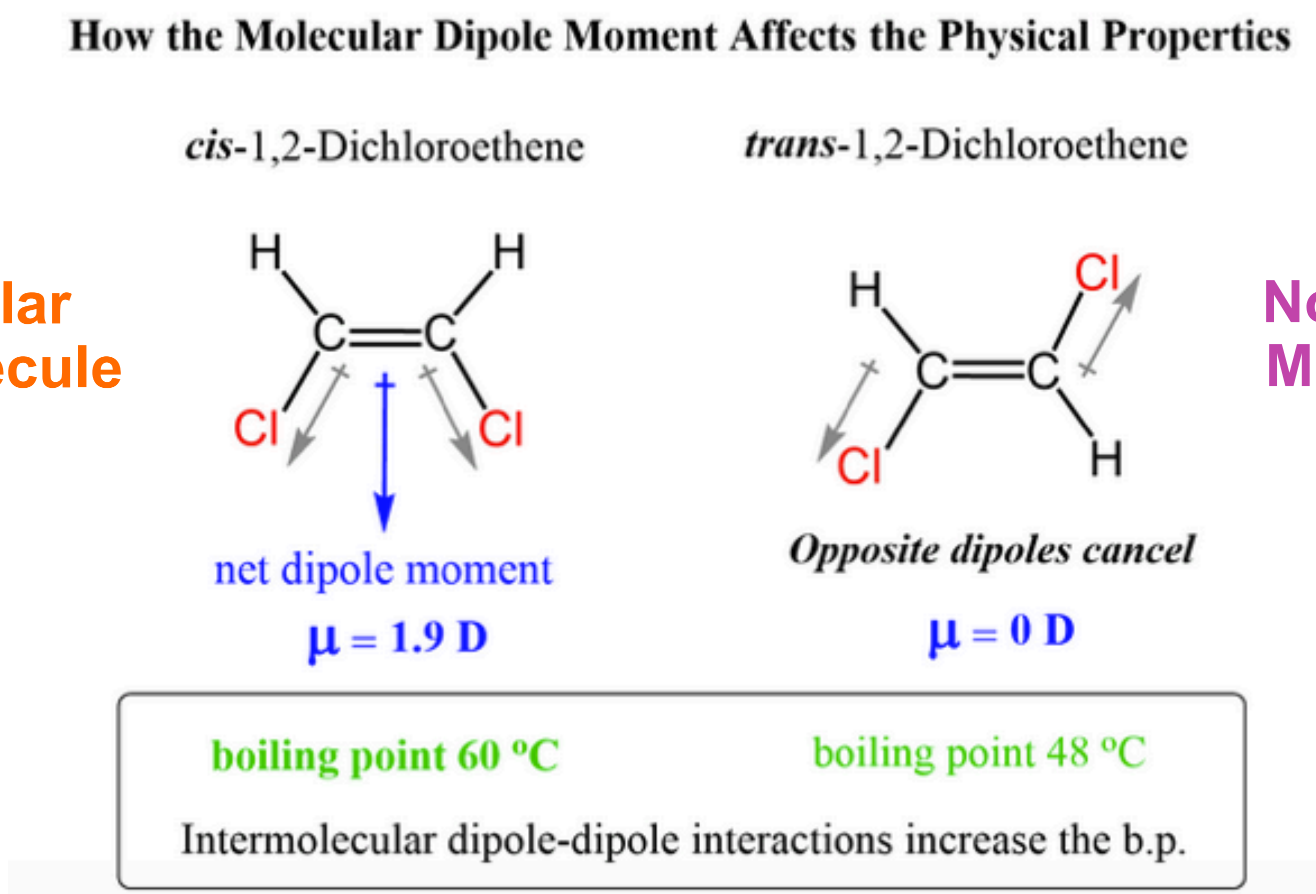


trans-1,2-Dichloroethene

Are there Polar Bonds? Is the Molecule Polar?

The dipole moments of the C-Cl bonds in ***cis***-1,2-Dichloroethene reinforce and the molecule **exhibits a net dipole.** In contrast, the ***trans*** isomer has **no molecular dipole** since the C-Cl bonds are at the opposite direction and their dipole effects are canceled:

Polar Molecule



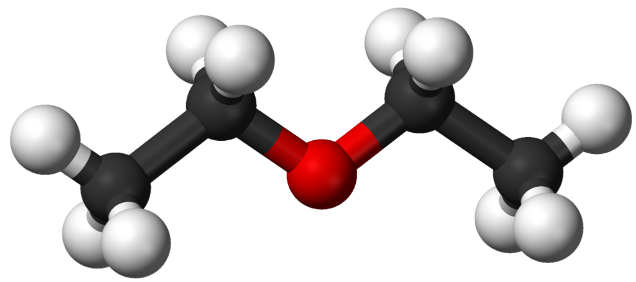
Nonpolar Molecule

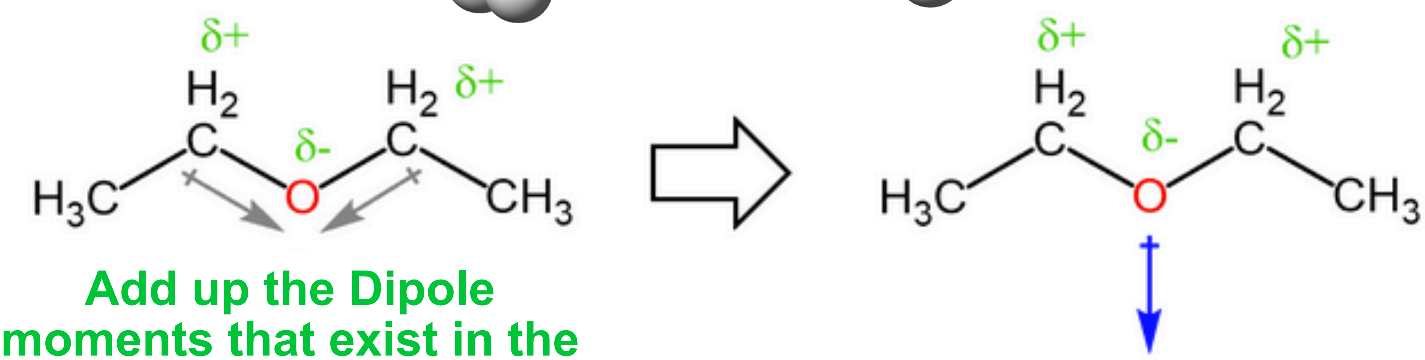
Are there Polar Bonds? Is the Molecule Polar?

Dipole Moment of Organic Molecules

For organic molecules containing a hydrocarbon chain with few polar bonds, the molecular dipole is determined by the sum of these dipole moments.

For example, diethyl ether possesses a net dipole moment which is the **vector sum of two C-O polar bonds**:

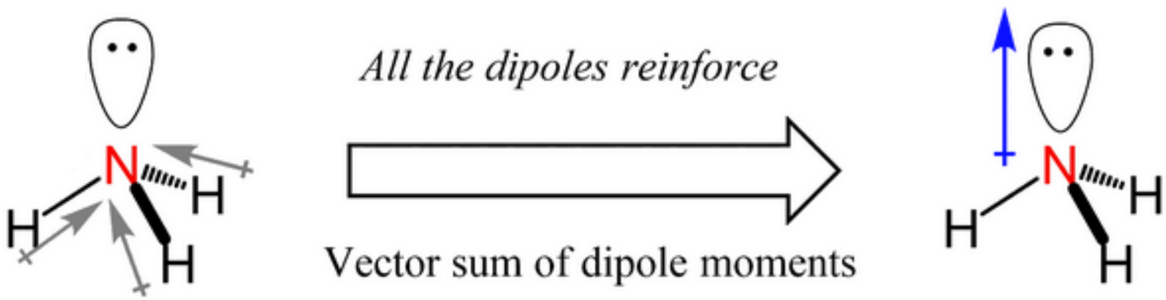




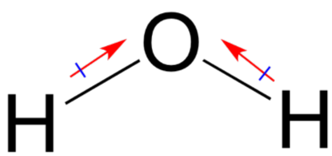
Add up the Dipole moments that exist in the bonds of the molecules

Forces do not cancel.
A net MOLECULAR dipole exists.
Molecule is polar.

Are there Polar Bonds? Is the Molecule Polar?

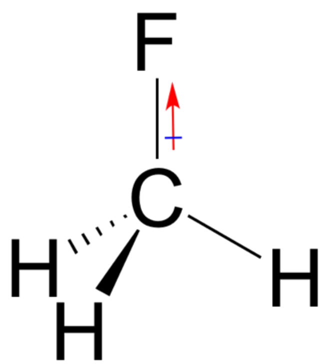


Molecule is polar.



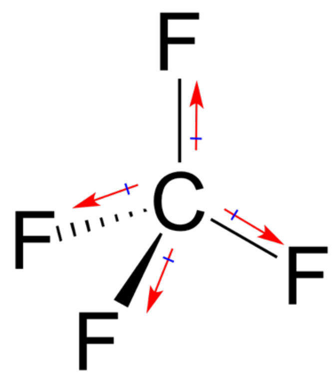
molecular dipole

Molecule is polar.



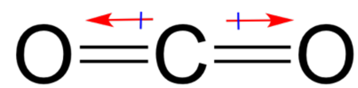
molecular dipole

Molecule is polar.



***no** molecular dipole*

Molecule is nonpolar.

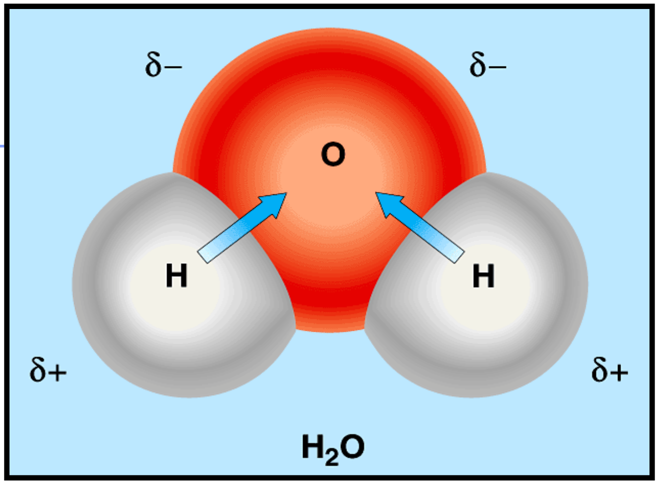


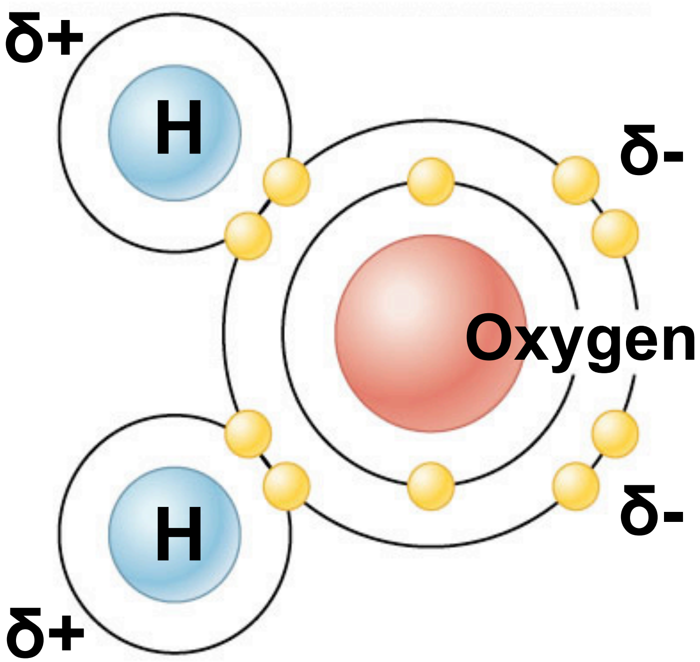
***no** molecular dipole*

Molecule is nonpolar.

Polar Molecules

- **Must have a pair of electrons in at least one or more bonds that are <u>shared unequally</u> by 2 atoms**
  - <u>Ex</u>: Water = H_2O
    - oxygen has stronger "attraction" for the electrons than hydrogen
    - oxygen has higher <u>electronegativity</u>
- <u>Polar Molecules</u>
  - Dipole moments (vectors existing in each polar bond) do <u>not</u> cancel each other out leaving a net directional force acting on the electrons in the molecule as a whole.
    - <u>Ex</u>: Water is a <u>polar molecule</u> (a dipole)
      - δ+ vs. δ– sides of the molecule exist
      - leads to many interesting properties of water…

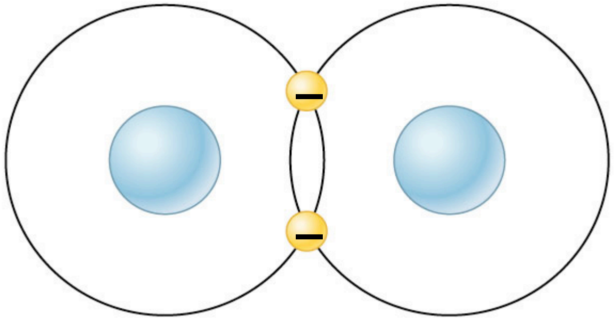




Biological Bonds

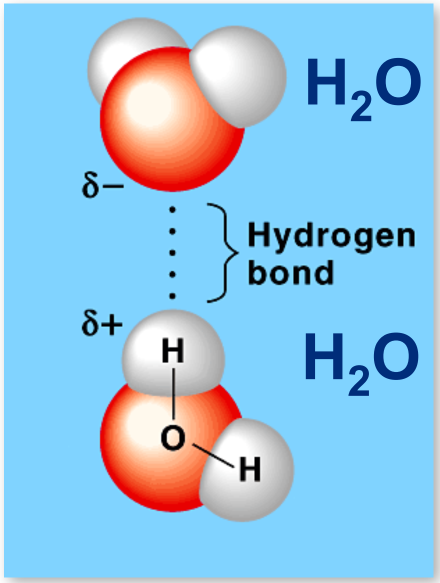
- Strong bonds (intramolecular)
  - Covalent bonds (bond can "only" be broken through a chemical reaction)
  - Ionic bond (bond can be broken through a chemical reaction or by dissociating when placed in water)
- Weak bonds (intermolecular attractions) are equally important!
  - Van der Waals Forces
    - Caused by attraction between DIPOLES (*i.e.* molecules with partial positive and negative 'ends' either permanent, temporary, or induced).
      - There are three sorts though intro-level chemistry/biology texts often erroneously equate Van Der Waals forces with London Dispersion Forces.
  - Ex: Hydrogen bonds
    - Attraction between a δ^+ H (covalently bonded to N, O, F) and a δ^- atom on another molecule)

Covalent bond

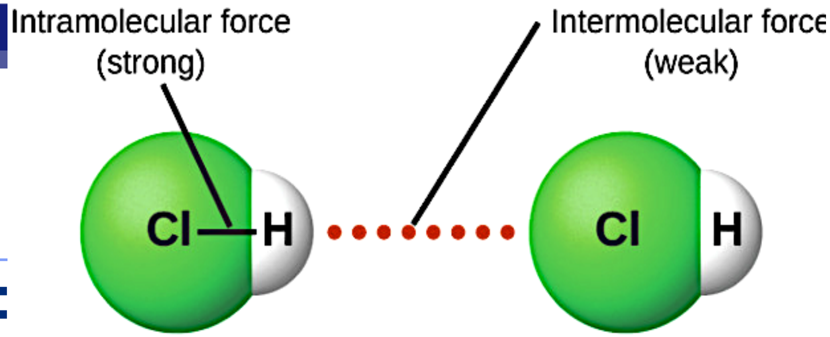


H_2 (hydrogen gas)

Hydrogen bond



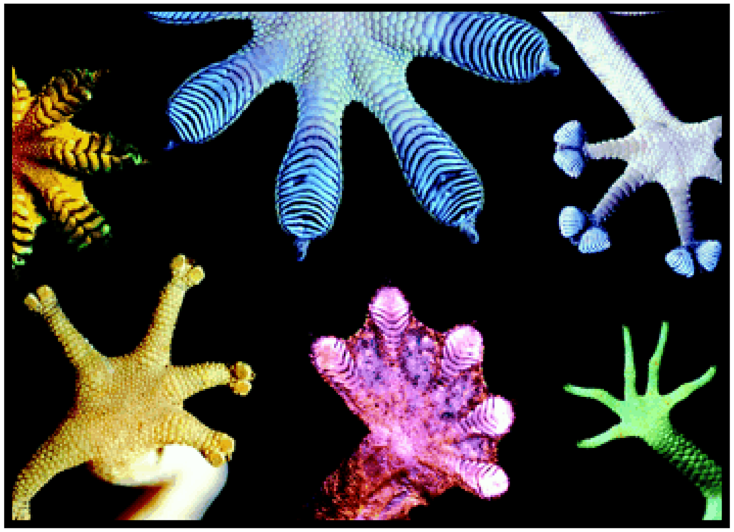
Biological Bonds



- During Physical Changes in matter:
  - Changing from solid, to liquid, to gas, for example
  - Molecular moving from one location to another, for example
    - Molecules are rearranged in space
    - Forces of attraction act between molecules
    - Intermolecular forces may be temporary
    - Intermolecular forces are broken as molecules move relative to one another
    - Intermolecular forces are weak
    - Interatomic (intramolecular) forces are not broken within each molecule
- During Chemical Changes in matter:
  - Covalent bonds are broken between atoms in reactants and atoms rearrange into new covalent bond partnerships in products
    - Intermolecular forces not involved
    - Atoms within a molecule are rearranged
    - Forces of attraction act within molecules
    - Intramolecular/interatomic forces may be persist for the life of the molecule (no matter in substance in solid, liquid, or gas state)
    - Intermolecular forces are strong
    - Interatomic (intramolecular) forces are are broken within each molecule

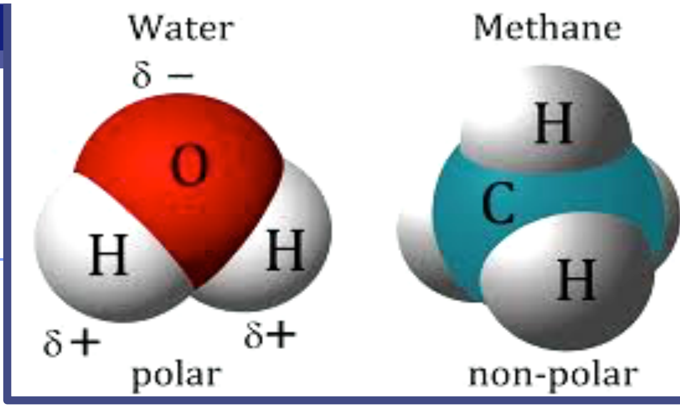
Van Der Waals Forces

- Electrons are not constantly symmetrically (evenly) distributed, even in nonpolar molecules and especially in polar molecules.
  - At times, they accumulate in one area or another even if only temporarily.
    - So there are longer- or shorter-lasting, yet ever-changing regions of + and - charges in molecules that can cause neighboring molecules to weakly attract each other at times.

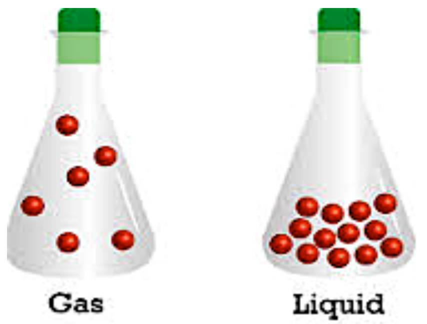


- **Types of Van Der Waals Forces:**
  1. Temporary Dipole - Induced Dipole Forces
     - A.K.A. London Dispersion Forces
  2. Dipole - Induced Dipole Forces
  3. Dipole - Dipole Forces
     - Including Hydrogen Bonding

Recall that Nonpolar molecules are not permanent dipoles



- Polar molecules have an imbalanced electron density distribution and thus charge distribution *- there are regions that are on average more permanently positively & negatively charged.*
- Nonpolar molecules have a more symmetrical electron density distribution.
  - They will be neutral and exhibit no net dipole moment.
    - Because nonpolar molecules lack any regions that display a more permanent net charge, non-polar molecules are not attracted easily to one another as well as to other polar and nonpolar molecules
      - Nonpolar molecules do not exhibit many intermolecular attractions.
        - Because there is always some thermal energy in a substance causing atoms to vibrate or molecules to collide randomly, nonpolar substances tend to form liquids and then gases at lower temperatures than polar substances of similar weight/size - there aren't any more permanent charges in the molecule that cause these nonpolar molecules to attract one another more regularly and hold them more closely together for longer in the presence of increasing kinetic energy as happens with polar substances.

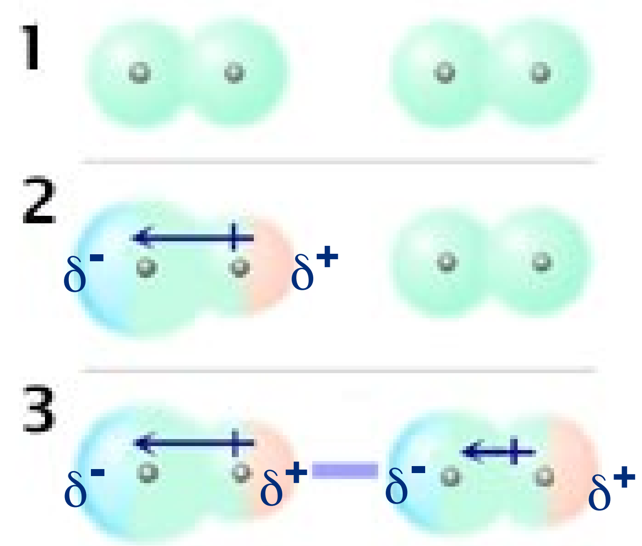


London Dispersion Forces
(the weakest of all Inermolecular Forces

These are attractions between ALL molecules (including NONPOLAR molecules).

Dispersion forces or London Dispersion forces are a result of electrostatic attraction between temporary & induced dipoles caused by the temporary variations in electron density around atoms and molecules.

1. The electron clouds and average e^- distribution of two non-polar molecules have a certain symmetry or 'evenness'.
2. However at any instant the electrons in an atom or molecule may be at one end, producing a temporary dipole moment.
3. This temporary dipole moment can induce a dipole moment in nearby molecules causing the two molecules to be attracted temporarily to each other.



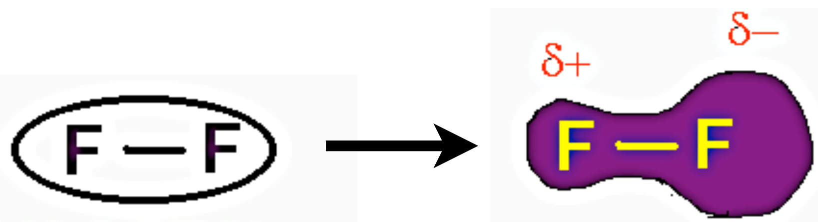
LONDON DISPERSION FORCES

EXAMPLE: **Consider the non-polar molecule F_2.**

Here, electrons are <u>shared equally</u> but are <u>constantly in motion</u>.

At times, the bonding electrons end up around one nuclei and not the other.

Thus, just for this instance, one end of the molecule becomes partially negatively charged, while the other becomes partially positively charged, forming a <u>TEMPORARY DIPOLE.</u>



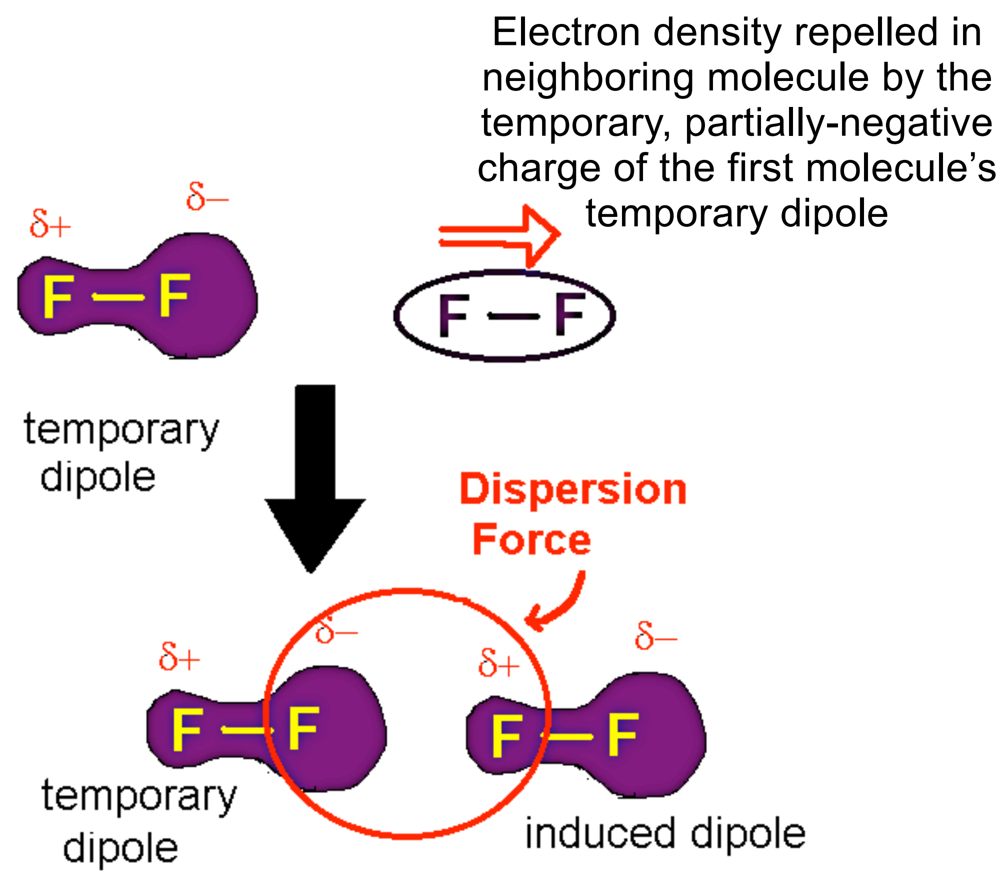
* on average, the electrons are shared equally between the two nuclei

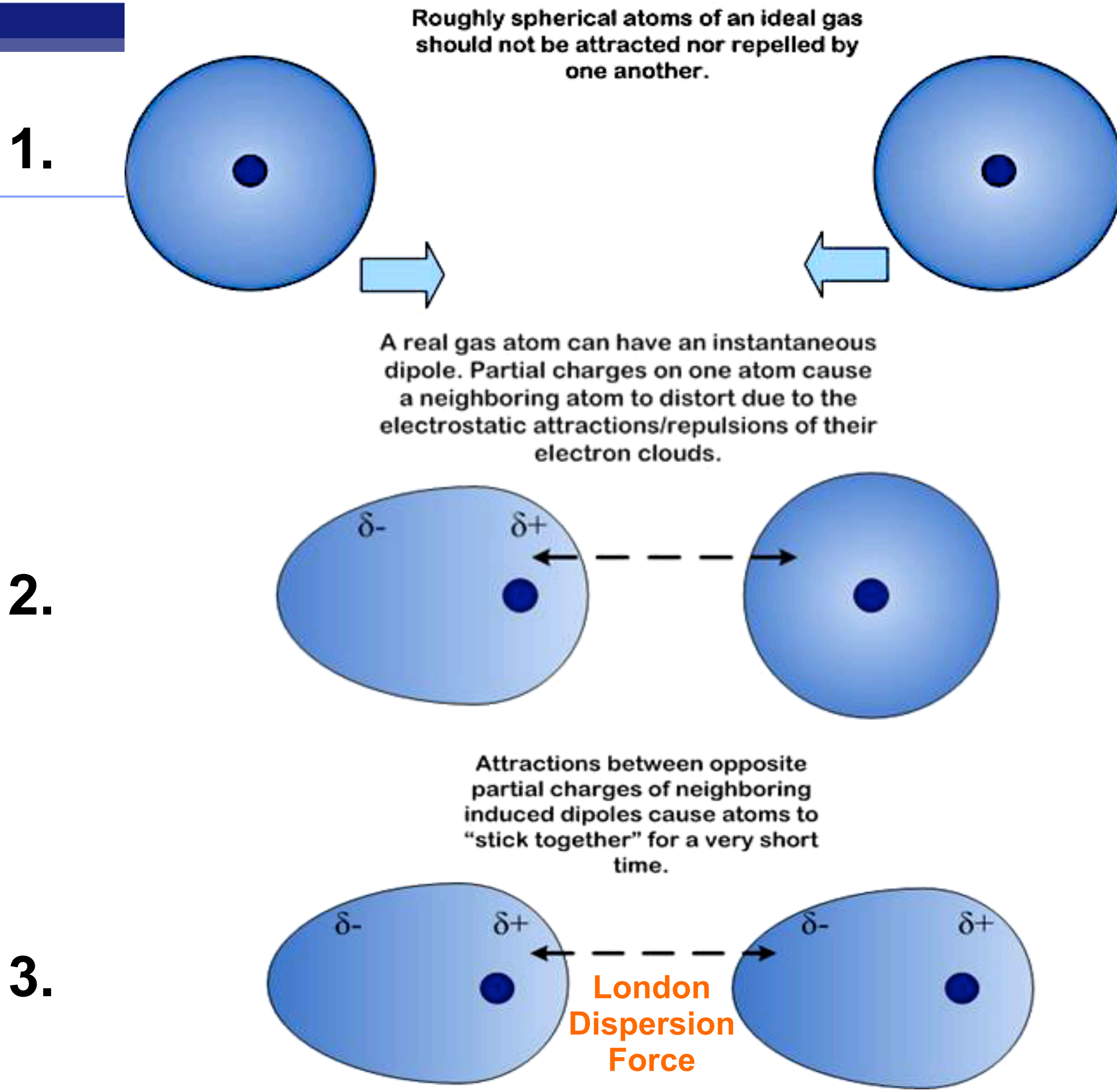
*occasionally, the shared electrons "spend more time" around one nucleus, making it partially negatively charge

LONDON DISPERSION FORCES

These temporary dipoles can attract or repel the electron density of nearby non-polar molecules, shifting the electron distribution in the neighboring molecule and producing a momentary induced dipole.

These dipoles may exist for only a fraction of a second, but a force of attraction between them also exist for that fraction of time.

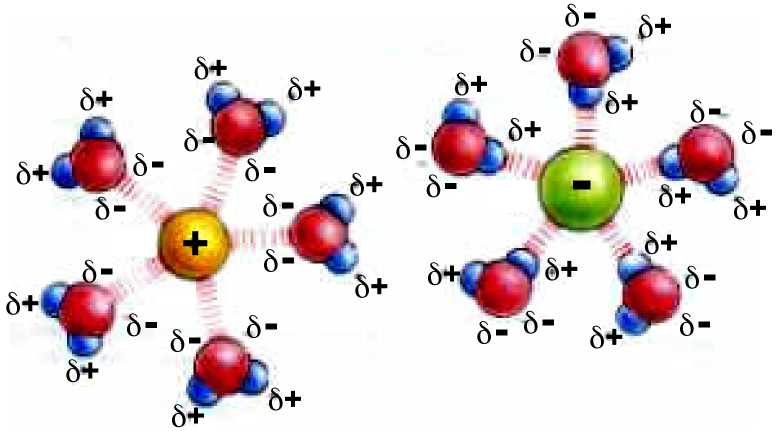



1.
Roughly spherical atoms of an ideal gas should not be attracted nor repelled by one another.
2.
A real gas atom can have an instantaneous dipole. Partial charges on one atom cause a neighboring atom to distort due to the electrostatic attractions/repulsions of their electron clouds.
δ-
δ+
3.
Attractions between opposite partial charges of neighboring induced dipoles cause atoms to "stick together" for a very short time.
δ-
δ+
δ-
δ+
London Dispersion Force

A quick review of the effects of Ions

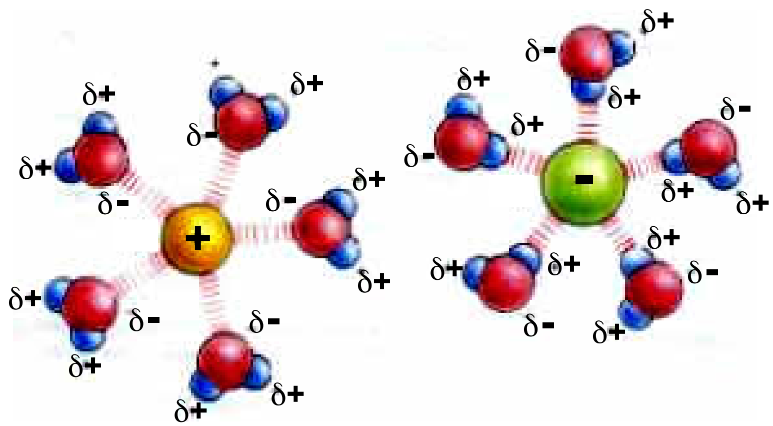
Before discussing (Permanent) Dipole - Induced Dipole forces and (Permanent) Dipole - (Permanent) Dipole Forces, let's review two other inter-molecular forces involving ions, intermolecular attractions that are stronger than the other intermolecular forces of attraction.

- **Ion - Dipole Forces:**
  - **When an ion nears a polar molecule, oppositely charged partial charges in the molecule are attracted to the ion.**

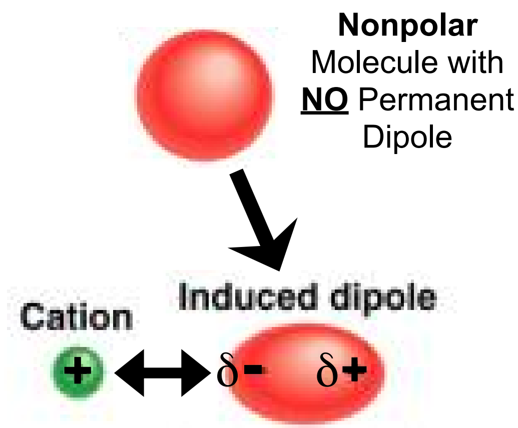


A quick review of the effects of Ions

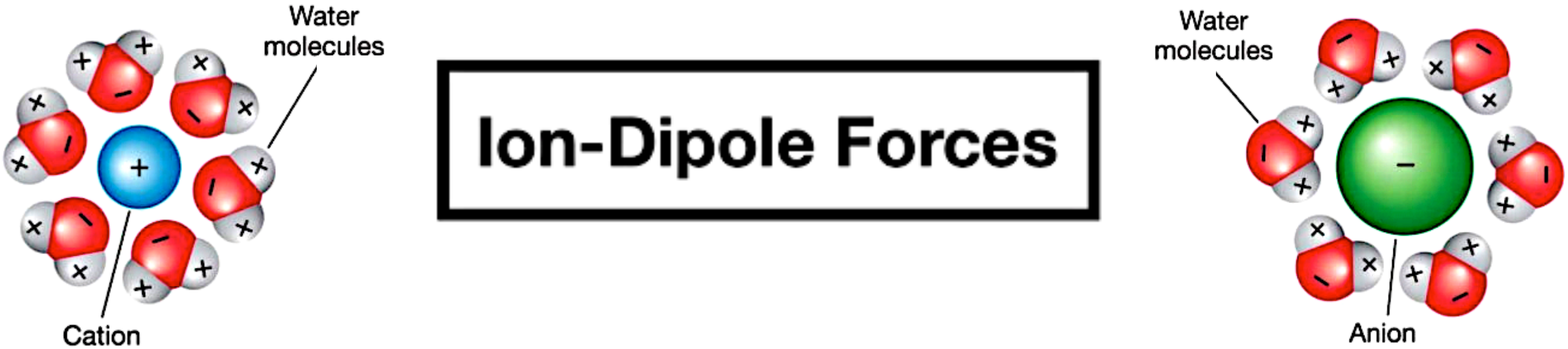
Before discussing (Permanent) Dipole - Induced Dipole forces and (Permanent) Dipole - (Permanent) Dipole Forces, let's review two other inter-molecular forces involving ions, intermolecular attractions that some of the strongest of the weak intermolecular forces of attraction.

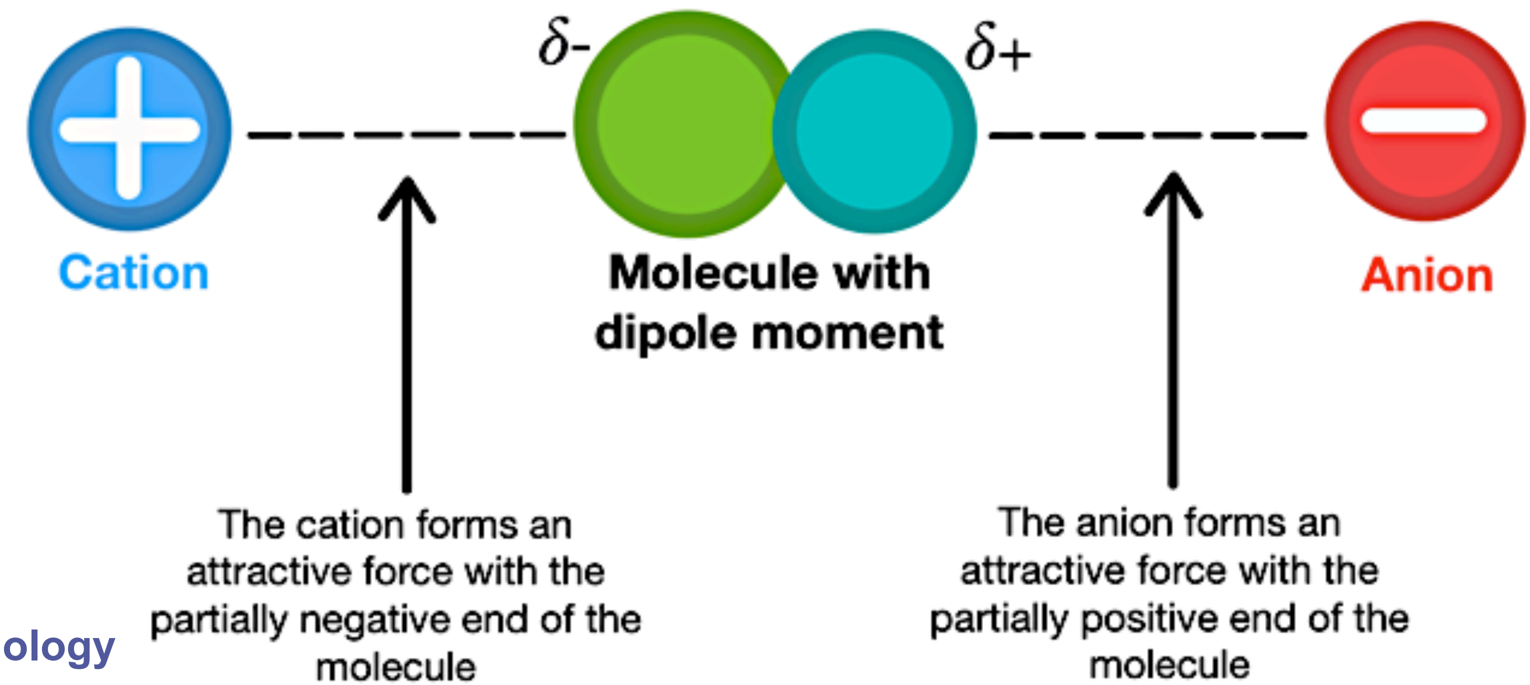


- **Ion - Dipole Forces:**
  - When an ion nears a polar molecule, oppositely charged partial charges in the molecule are attracted to the ion.
- **Ion-Induced Dipole Force:**
  - Weak attraction results when an ion induces a temporary dipole in a nonpolar molecule by disturbing the arrangement of electrons in the nonpolar species.

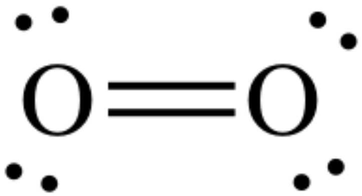


A quick review of the effects of Ions

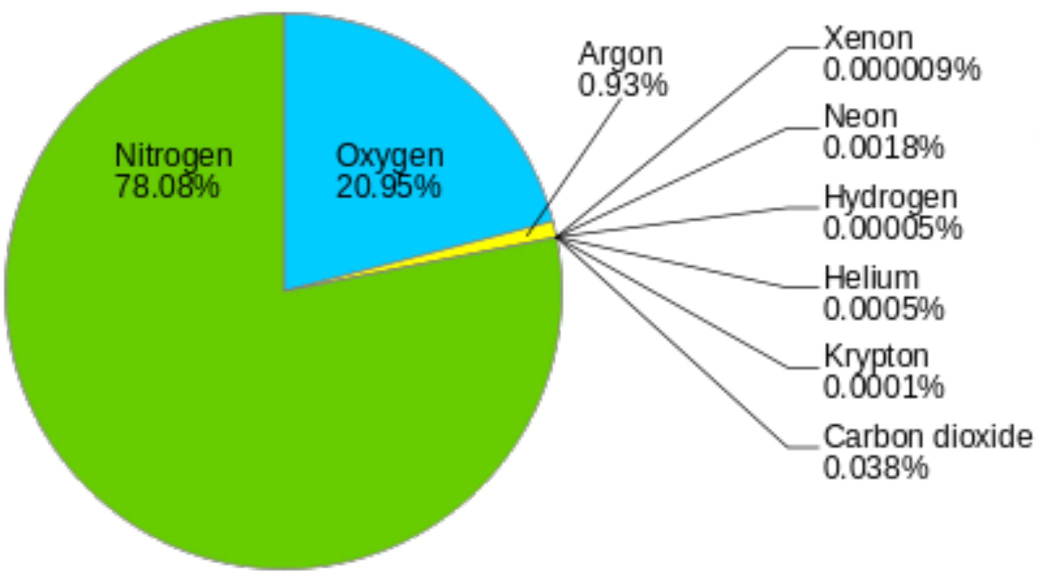


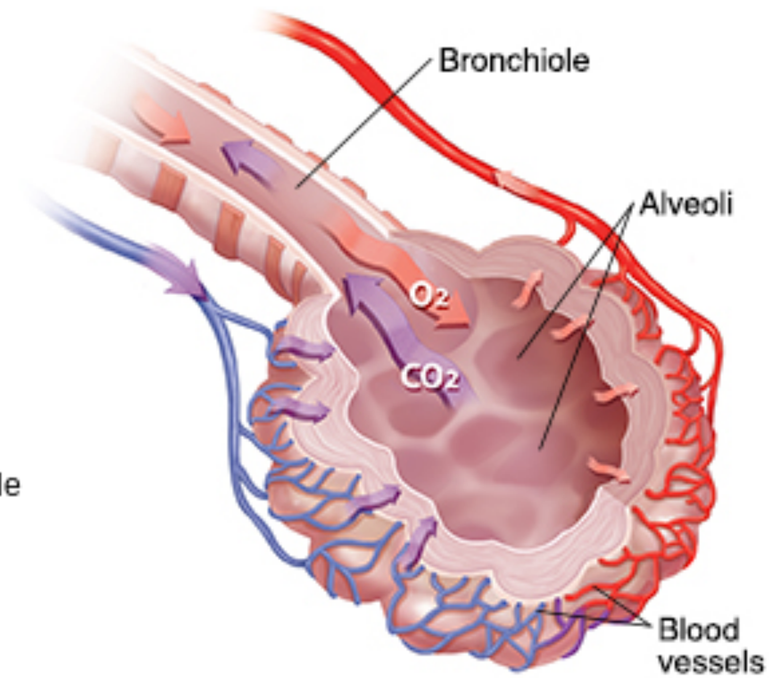


Ion - Induced Dipole Forces

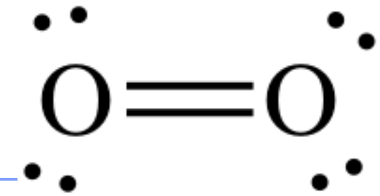


- How do humans carry oxygen to the cells of the body?
  - Oxygen (O2) is a non-polar molecule
    - It has <u>no permanent dipole moment</u> given that both O atoms are equally electronegative.
  - Air is filled with 21% O2 gas.
    - We inhale air bringing the air into our lung cavity in order to bring O2 gas close to blood vessels that surround our lung tissue

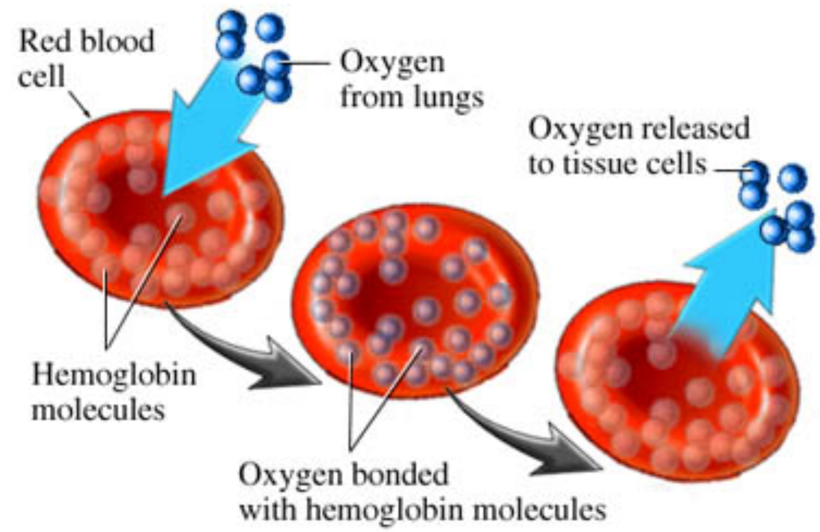




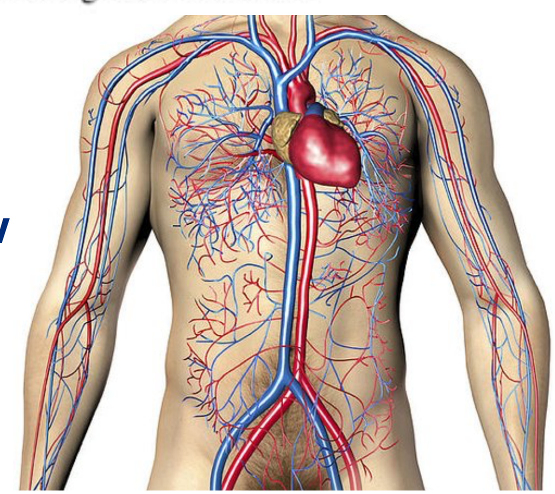
Ion - Induced Dipole Forces



- It's the job of red blood cells to carry oxygen to all cells
  - Red blood cells are filled with a transport protein known as hemoglobin

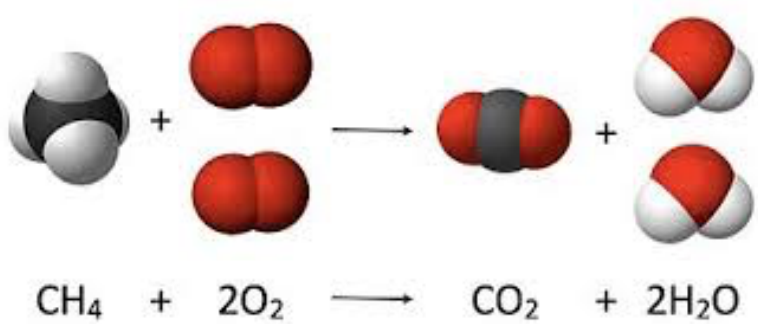


    - When you inhale air high in oxygen content, oxygen molecules diffuse through the lung into the surrounding blood vessels, diffusing into red blood cells where they attach to hemoglobin proteins.
    - As the heart pumps blood through the body, red blood cells carry oxygen throughout the body
    - The oxygen molecules will detach from hemoglobin when they reaches body cells low in oxygen, the oxygen gas diffusing out of the red blood cell, out of the blood vessel, and into the body cells.



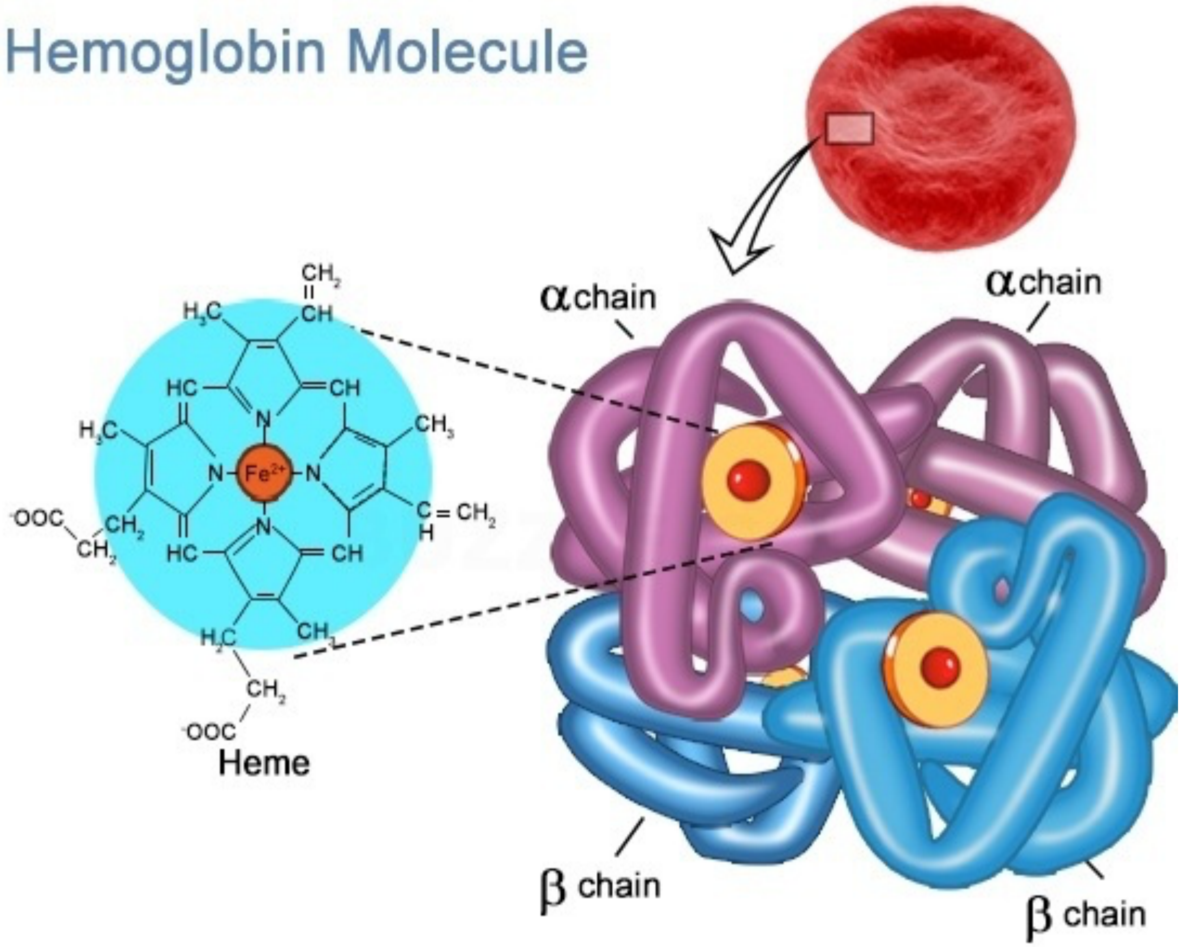
Ion - Induced Dipole Forces

- To be efficient, oxygen must quickly bind to hemoglobin in the lung and yet also easily detach from hemoglobin when the red blood cell reaches tissues, ideally without requiring a lot of effort or energy.
  - **Recall:** Chemical reactions involve breaking strong intramolecular forces (ionic & covalent bonds) in order to rearrange atoms of reactants into new products.
    - Breaking covalent and ionic bonds between atoms isn't easy - it takes an input of energy or the use of additional substances, called catalysts.
      - If oxygen bound to hemoglobin through covalent or strong ionic bonds, it would be hard to release the oxygen when red blood cells reach body tissues.
        - Oxygen thus bonds through a weaker intermolecular force: an Ion - Induced Dipole Force



Ion - Induced Dipole Forces

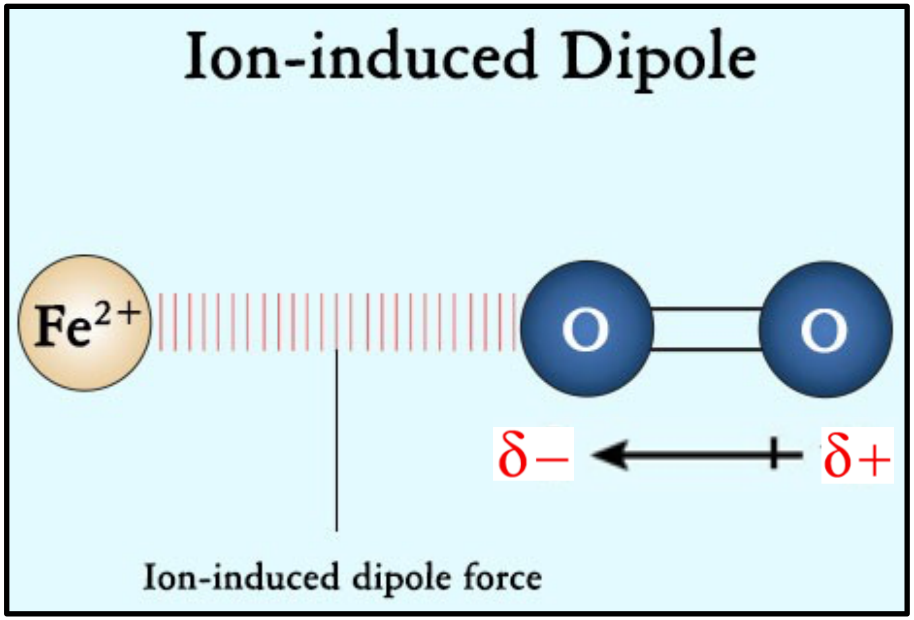
- Hemoglobin is a protein made up of four large molecules or subunits, <u>two alpha chains and two beta chains</u>.



  - At the center of each of the 4 subunits or chains, you will find an <u>Iron Cation</u>

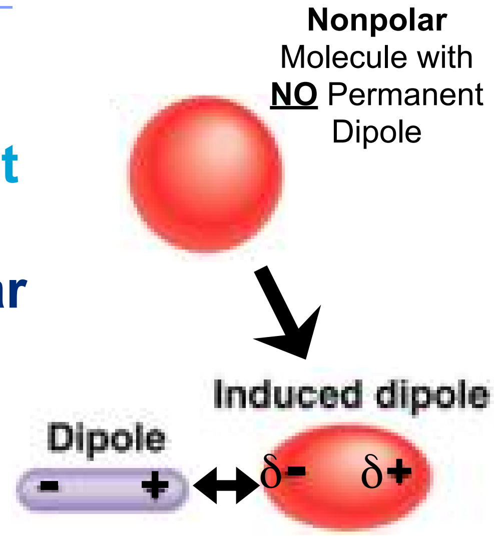
Ion - Induced Dipole Forces

- When O_2 molecules diffuse into the red blood cells and pass near the 4 Fe2+ ions on each hemoglobin transport protein, the strong positive charge of the Fe2+ cation, causes an <u>induced dipole to form in the O_2 molecule.</u>
  - The O_2 molecule now experiences a weak intermolecular force of attraction between it and the Fe2+ ion.
    - The resulting ion-induced dipole force holds O_2 onto the hemoglobin.
  - Since this force of attraction is much weaker than any strong covalent/ionic bond, O2 can easily detach from hemoglobin in order to diffuse into body cells low in oxygen as blood in pumped through body.

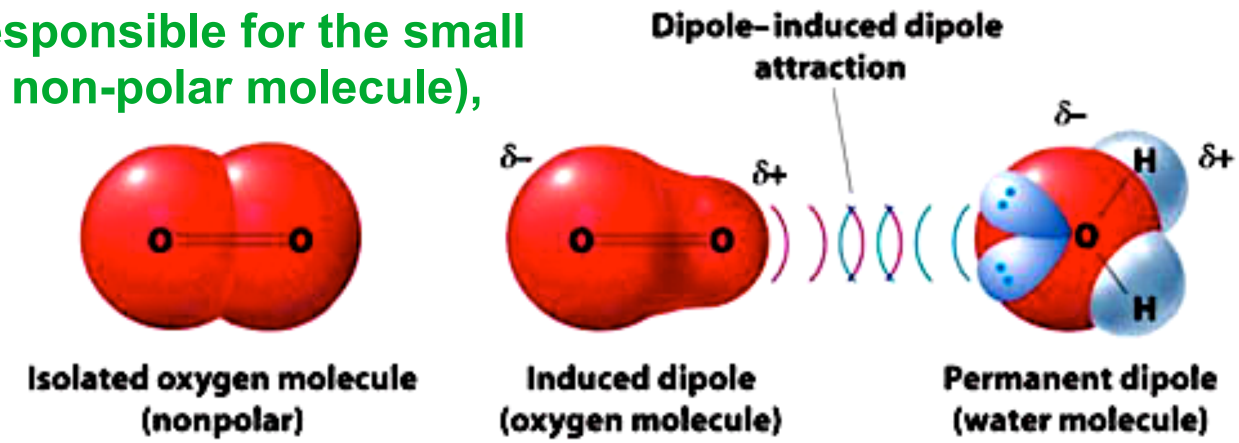


Permanent Dipole - Induced Dipole Forces

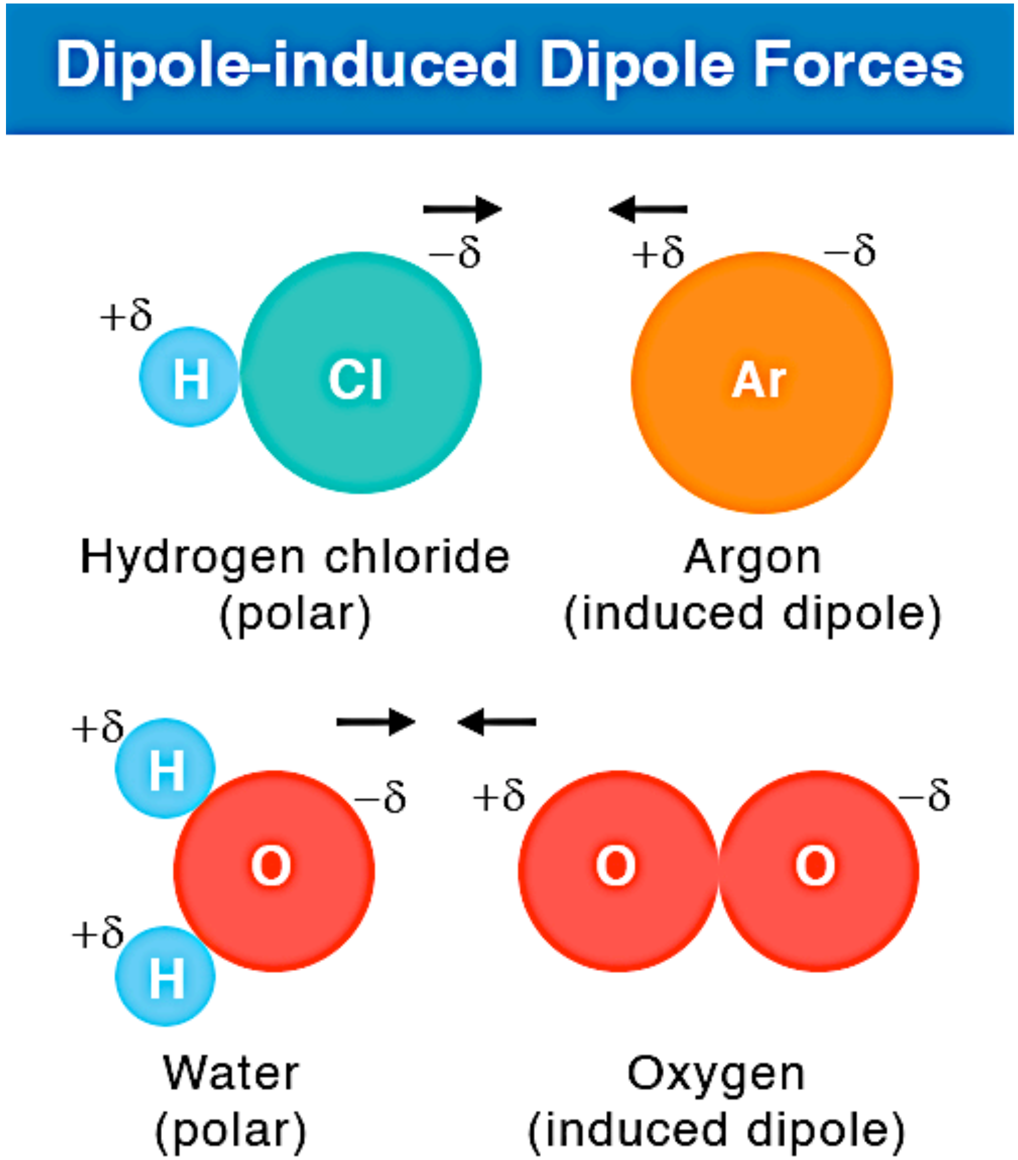
- **Dipole - Induced Dipole Forces:**
  - A polar molecule (a molecule with a permanent dipole) produces an electric field which can distort the electron cloud of a nearby non-polar molecule, inducing a dipole.
  - This will cause the an attraction between the molecules
  - When the induced dipole moves away, it looses its temporary polarity going back to a non-polar molecule.



This type of force is responsible for the small solubility of oxygen (a non-polar molecule), known as dissolved oxygen, in water (a permanent dipole).

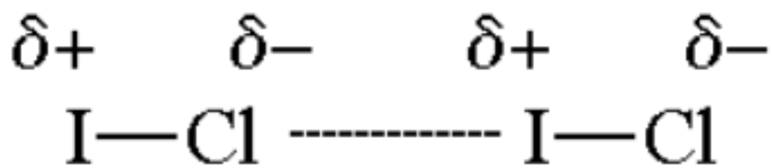


Permanent Dipole - Induced Dipole Forces

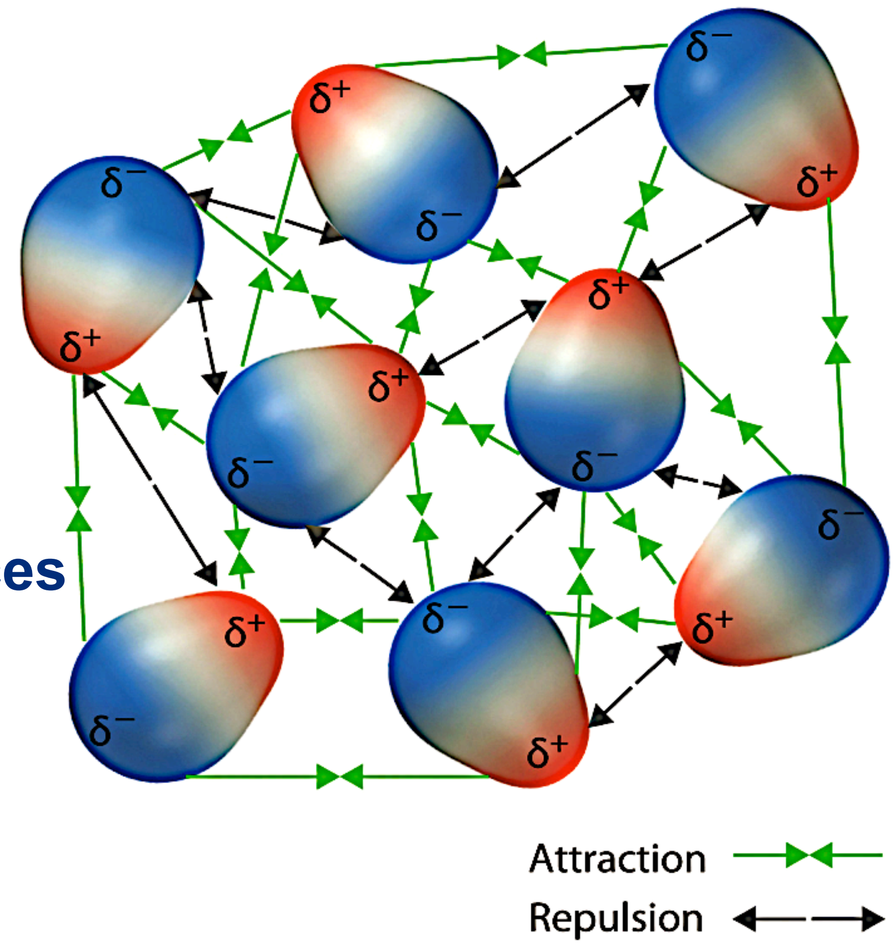


(Permanent) Dipole-Dipole Forces

If two neutral <u>POLAR molecules</u>, each being permanent dipoles, come together such that their <u>oppositely charged ends align</u>, they will be <u>attracted</u> to each other.



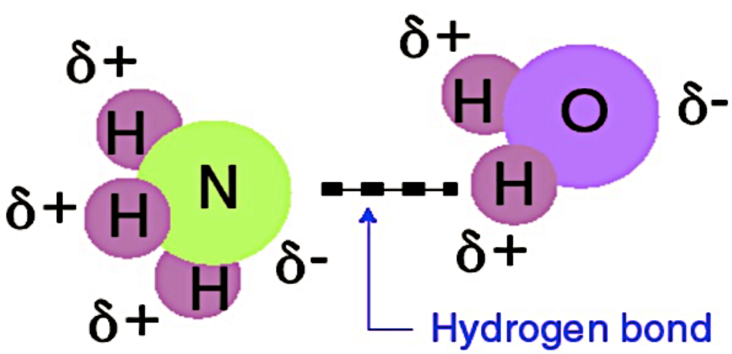
Dipole-dipole attractions are stronger than Dipole-Induced Dipole forces, the latter being stronger than Lond. Disp. Forces

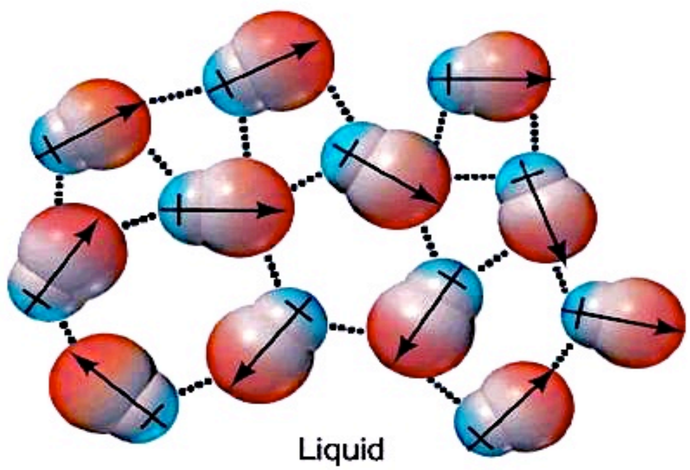


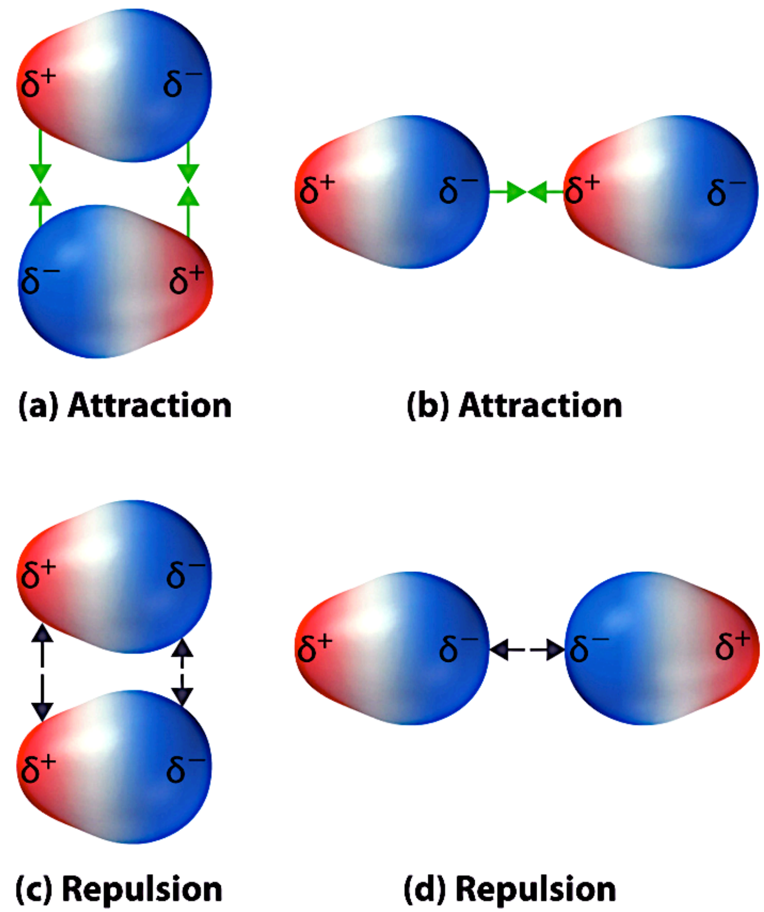
A special type and one of the strongest of the weak dipole-dipole interaction is a <u>hydrogen bond</u>

(Permanent) Dipole-Dipole Forces

Polar molecules arrange themselves (rotate) in ways that maximize the attractions between opposite charges and minimize the repulsions between like charges.

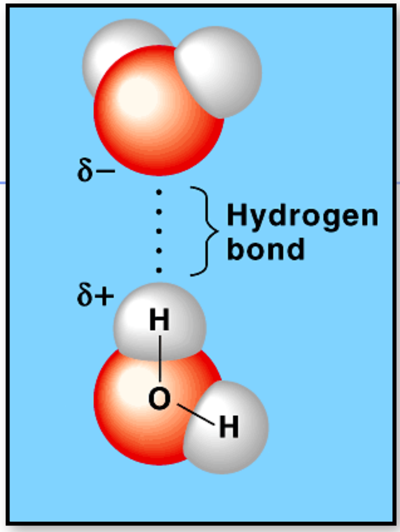


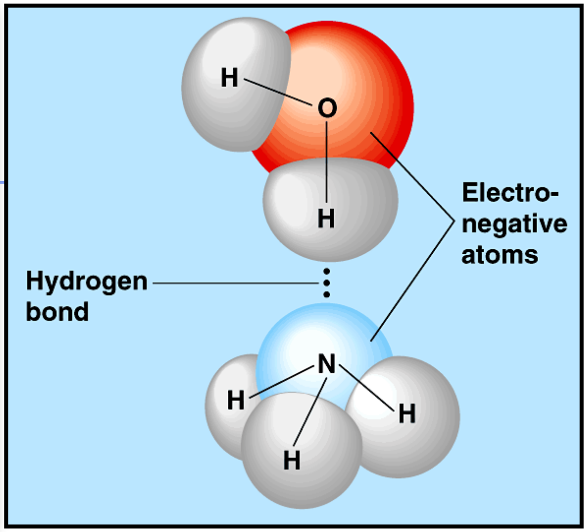




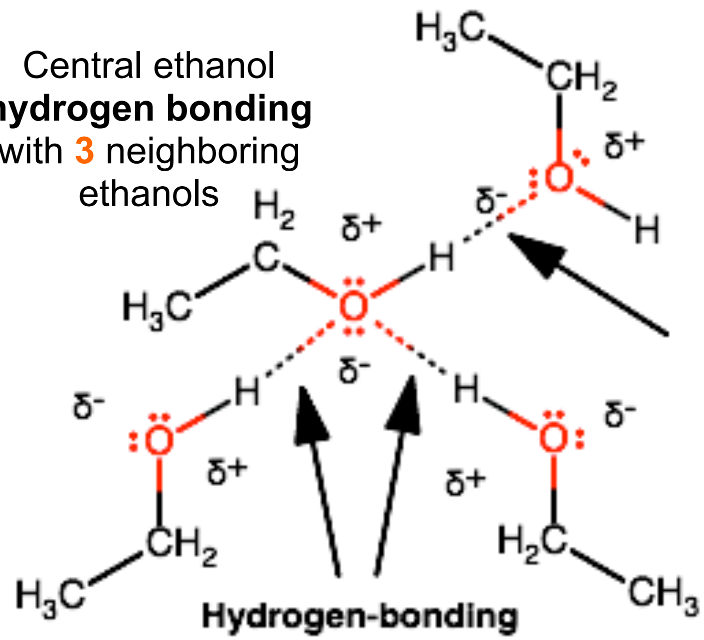
Hydrogen bonding -

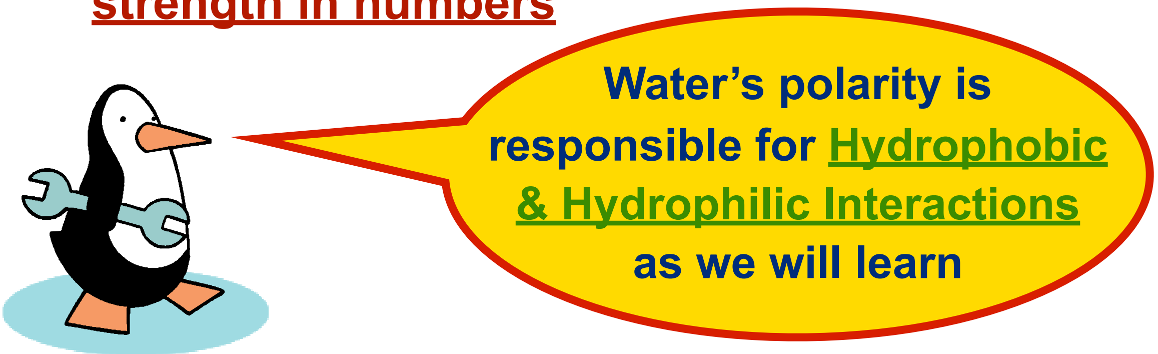
A special case of dipole-dipole interaction





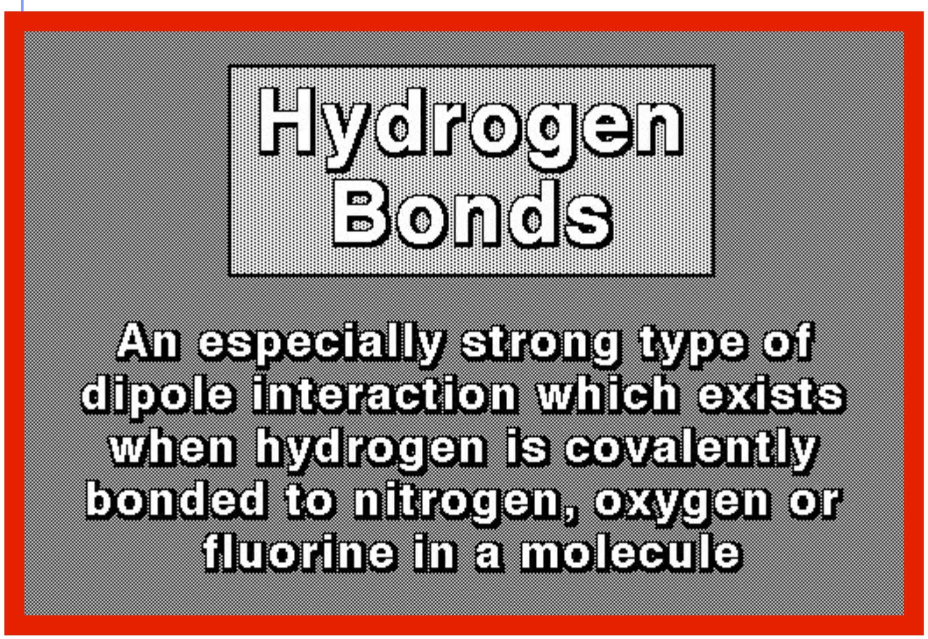
- The polarity of water results in intermolecular attractions
  - Attraction between a partially positive H in one H_2O molecule to a partially negative O in another H_2O
  - Can also occur wherever an -OH or -NH group exists in a larger molecule
- <u>Weak bond</u> (10x weaker than covalent or ionic)
- Bonds can be broken but there is <u>strength in numbers</u>

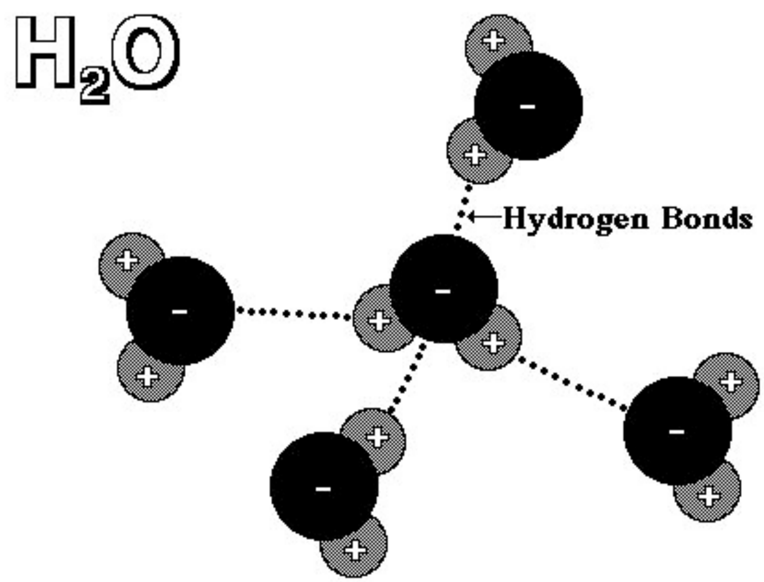




HYDROGEN BONDING

Hydrogen Bonding = A special type of dipole force (permanent dipole – permanent dipole interaction)





HYDROGEN BONDING

Hydrogen Bonds are about 15 to 100 times weaker than covalent bonds BUT they are the strongest type of dipole interaction (van der Waals force) – about 10x stronger than normal dipole forces (between polar molecules) for two reasons:

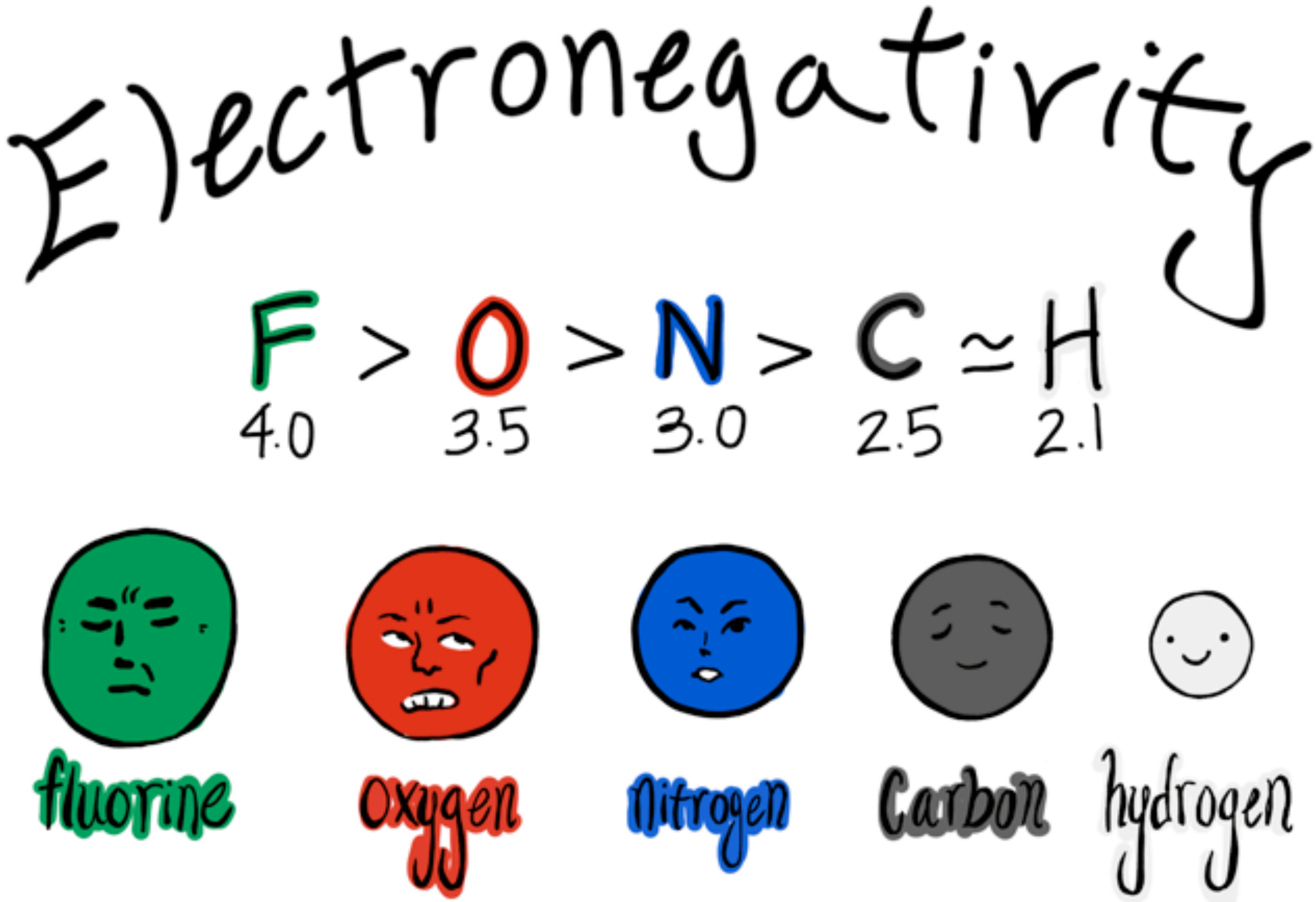
Hydrogen Bonds

Nitrogen, oxygen and fluorine are all very electronegative, so a covalent bond between any of them and hydrogen tends to be very polarized

Hydrogen Bonds

Also, hydrogen, nitrogen, oxygen and fluorine are small atoms, so the negative end of one dipole can approach very close to the positive end of another dipole, and vice versa

Relative Electronegativities



HYDROGEN BONDING

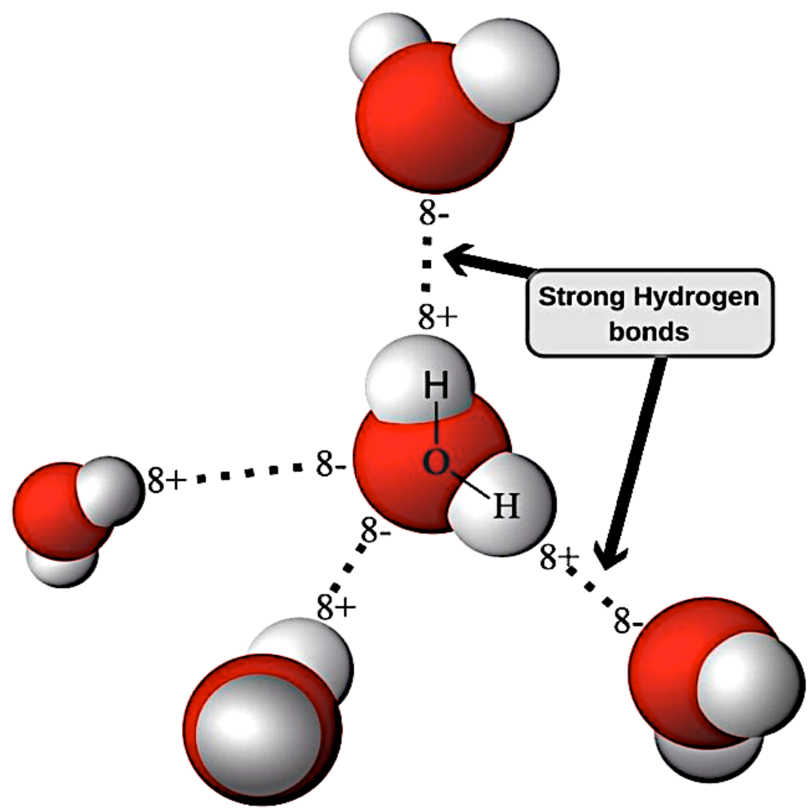
Why do Hydrogen Bonds form?

Because of the high electronegativity of O, N, and F, the electrons shared in the COVALENT BOND within a molecule between O, N, and F and H are more strongly attracted to the O, N, or F atom than the H atom.

The shared electrons "spend more time" in the electron cloud of O, N, or F compared to the H, forming a VERY polar bond.

The O, N, and F atoms become partially negatively charged, more so that in just any other polar molecule, but still they do not gain a full negative charge (the shared electrons do spend some "time" around the H nucleus).

Between molecules, now the δ^+ H in one molecule finds itself attracted to an δ^- O, N, or F atom in a neighboring molecule (the hydrogen bond)

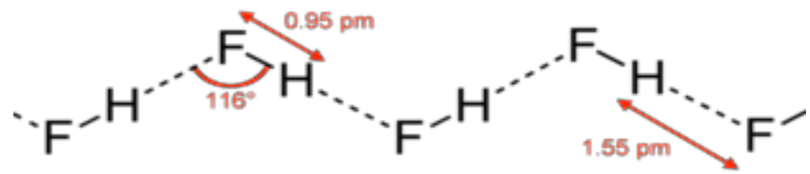


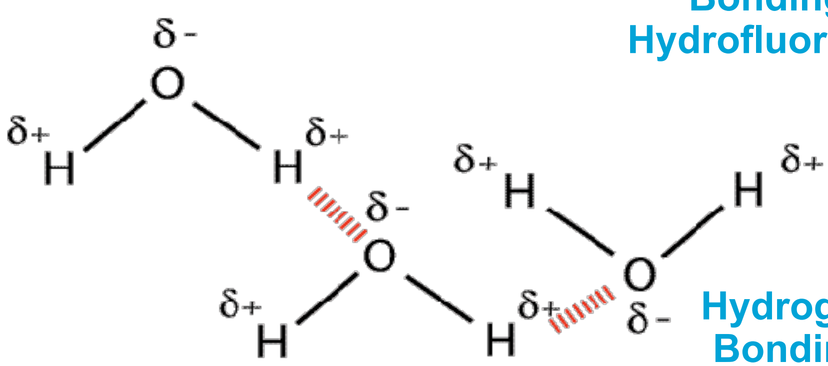
HYDROGEN BONDING

A hydrogen bond is formed when the negative or partially negative atom of one molecule becomes oriented to and attracted to the partially positive H of another molecule.

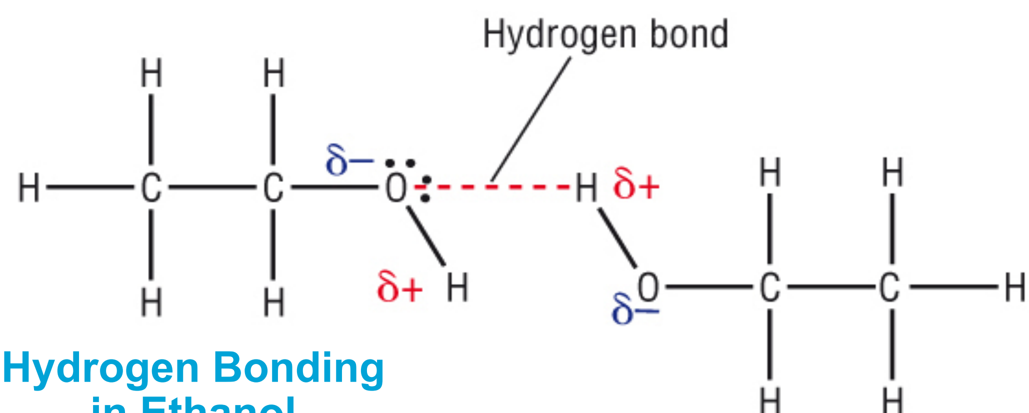
(H-bonds do not just form between water molecules!)

Hydrogen Bonding in Hydrofluoric Acid

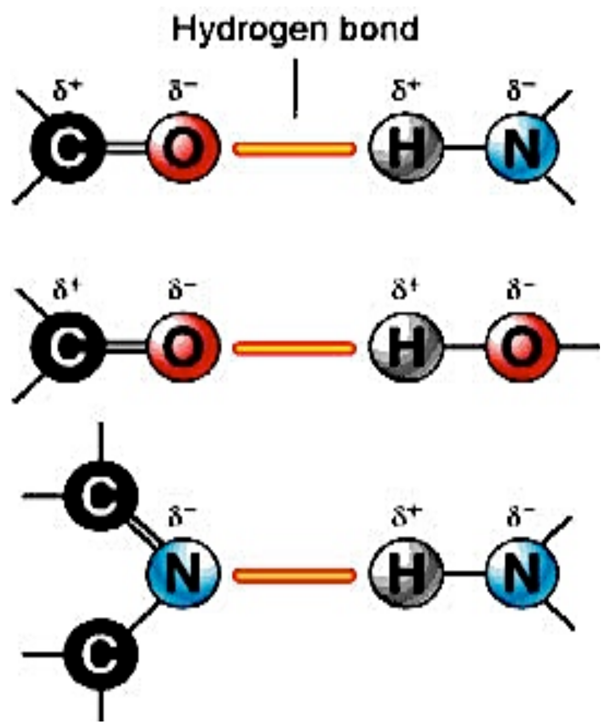




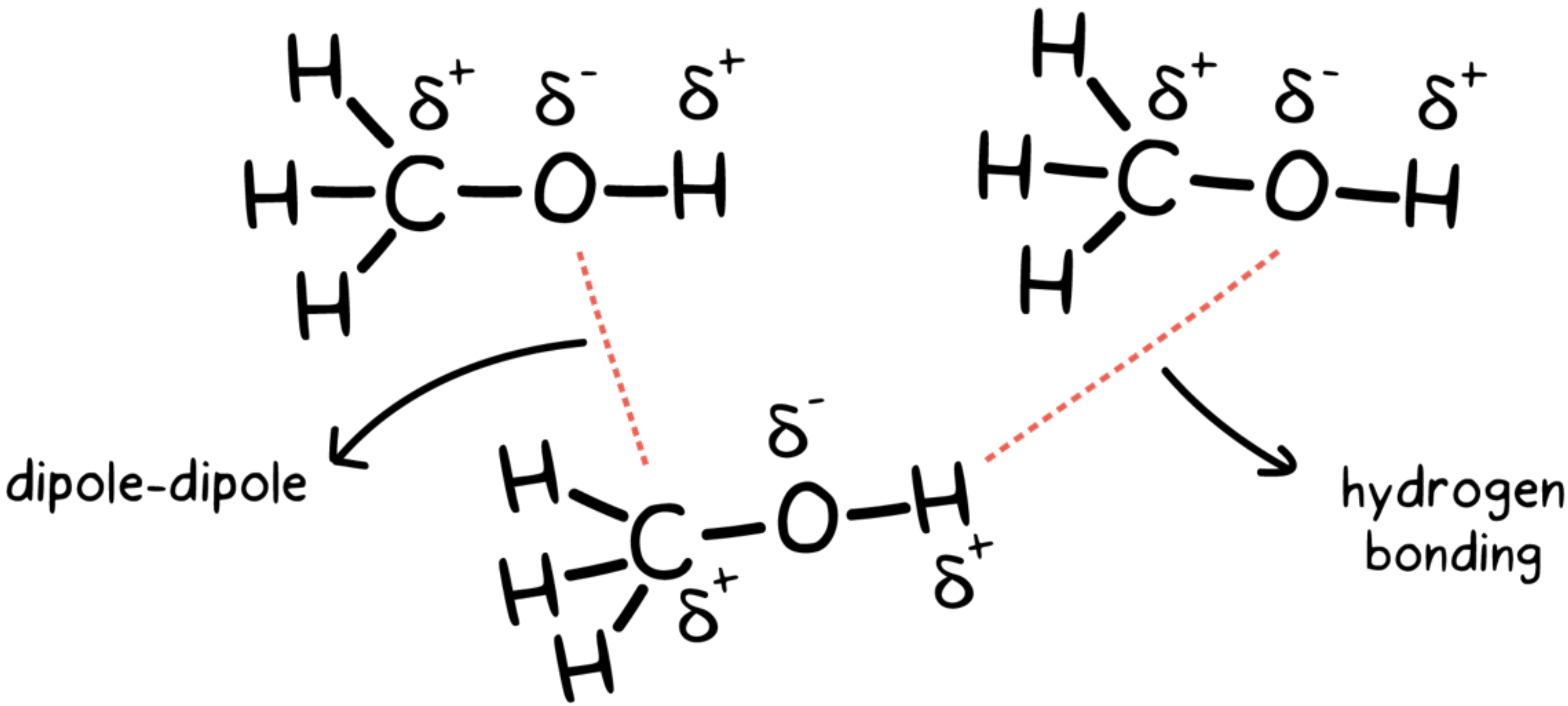
Hydrogen Bonding in Water



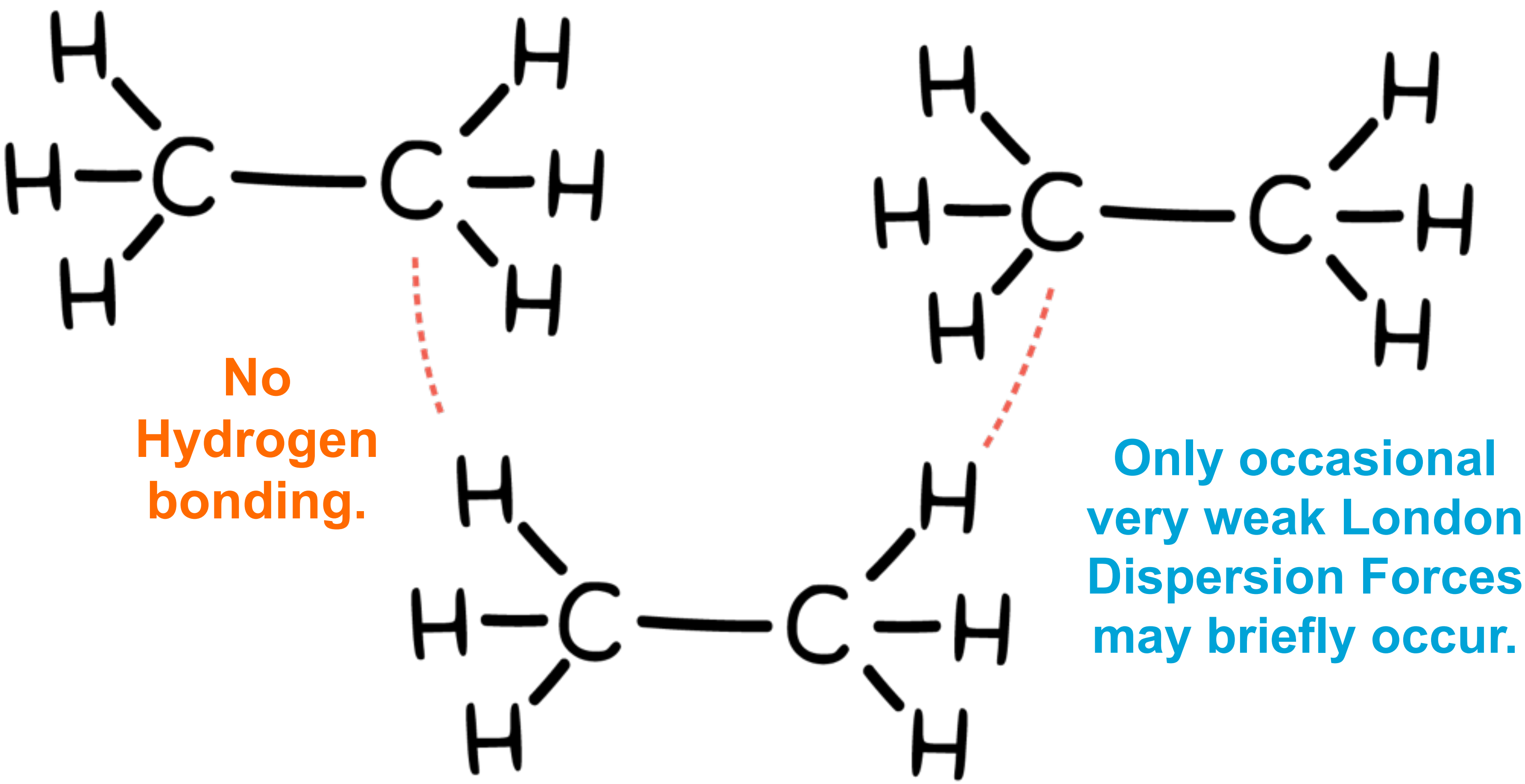
Hydrogen Bonding in Ethanol

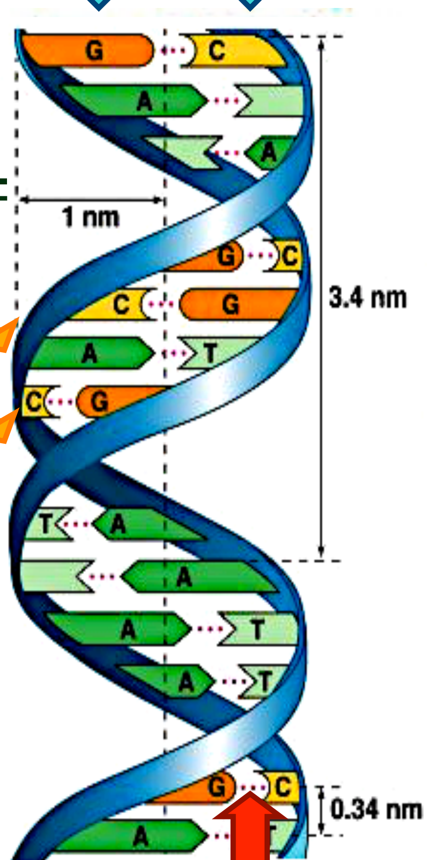


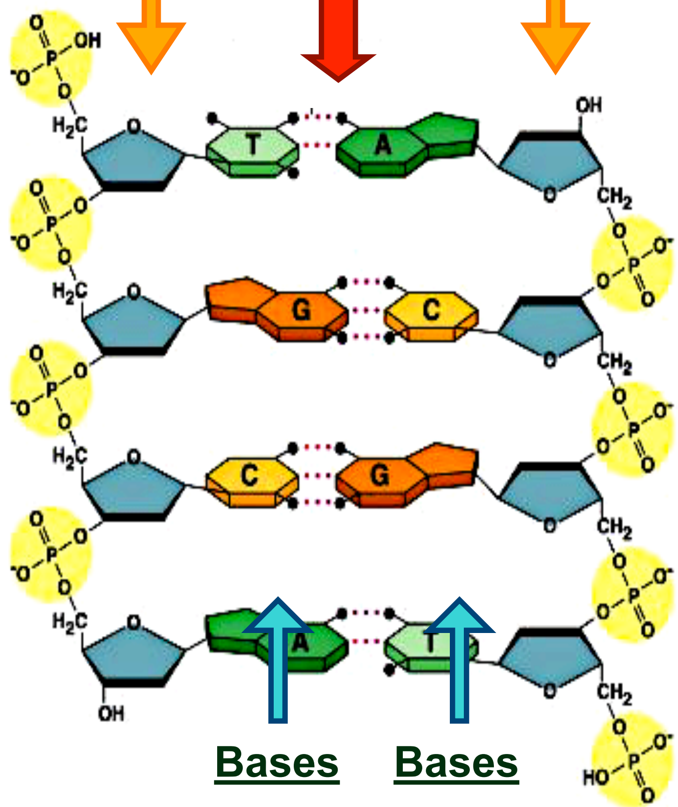
HYDROGEN BONDING OR NO HYDROGEN BONDING?



HYDROGEN BONDING OR NO HYDROGEN BONDING?

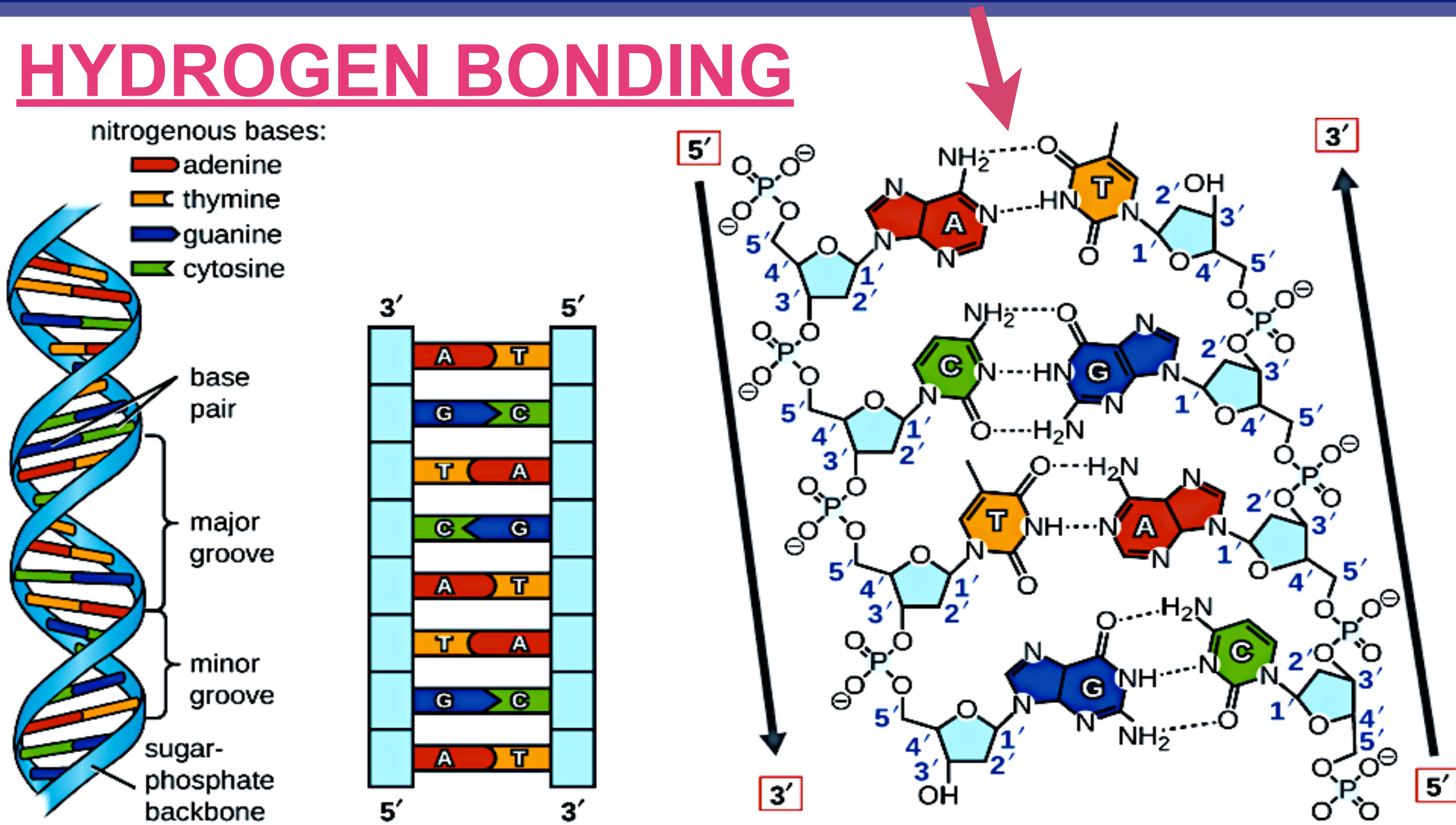






These 2 DNA strands are held together by
HYDROGEN BONDS
between nearby atoms in DNA molecules.

HYDROGEN BONDING



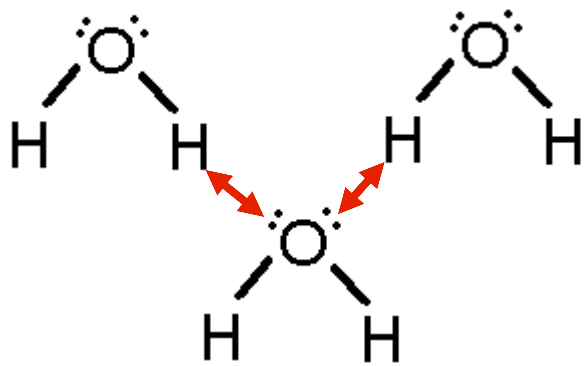
Hydrogen bonds are one the <u>strongest of the intermolecular forces</u> (besides ion-dipole and ion-induced dipole forces), but still <u>weak</u> when compared to a fully ionic (full + and - charge) attraction or compared to a covalent bond.

Even so, <u>given great numbers, Hydrogen Bonds can be</u> <u>quite strong</u> and provide a significant level of stabilization to large molecules such as proteins and nucleic acids (DNA or RNA).

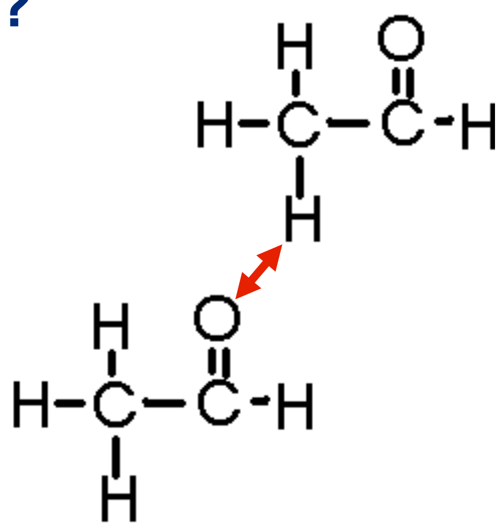
HYDROGEN BONDING OR NO HYDROGEN BONDING?

In order for hydrogen bonds to form between molecules, hydrogen MUST be bonded to nitrogen, oxygen or fluorine in the SAME molecule. ***The hydrogen bonds though occur BETWEEN molecules!!!***

Look at the following groups of molecules: Which would qualify as having hydrogen bonds?

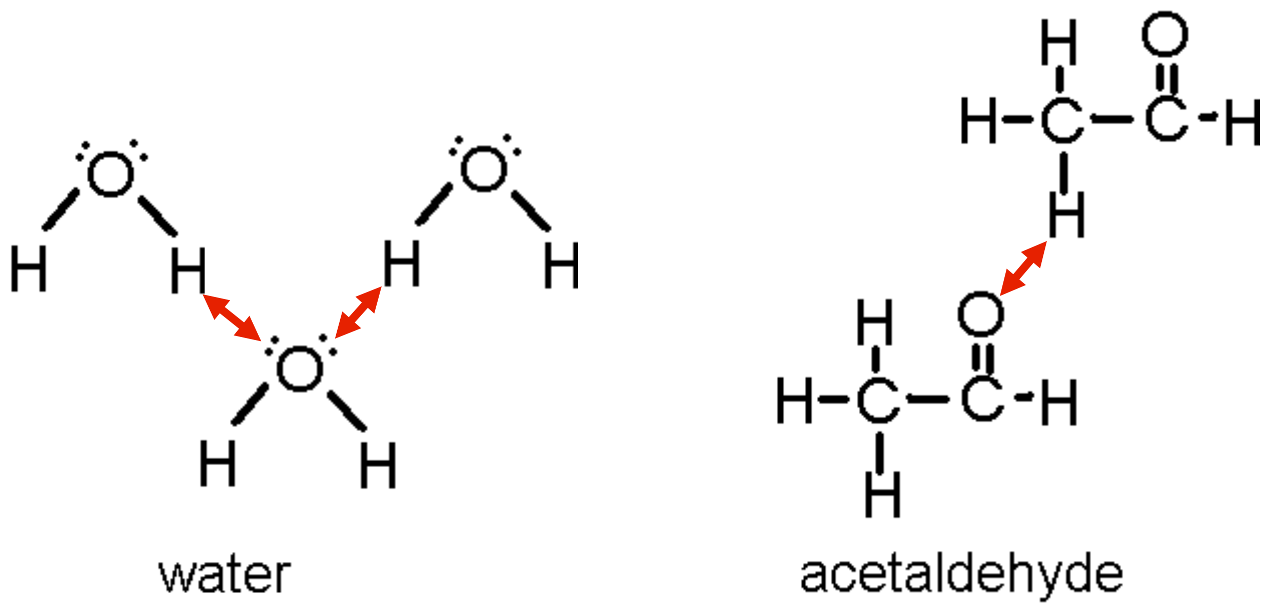


water



acetaldehyde

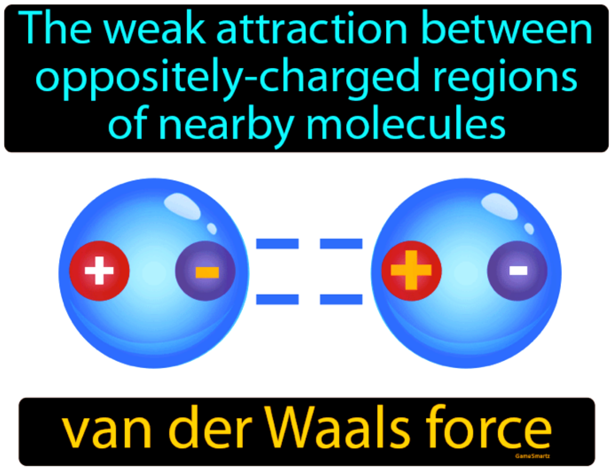
HYDROGEN BONDING OR NO HYDROGEN BONDING?

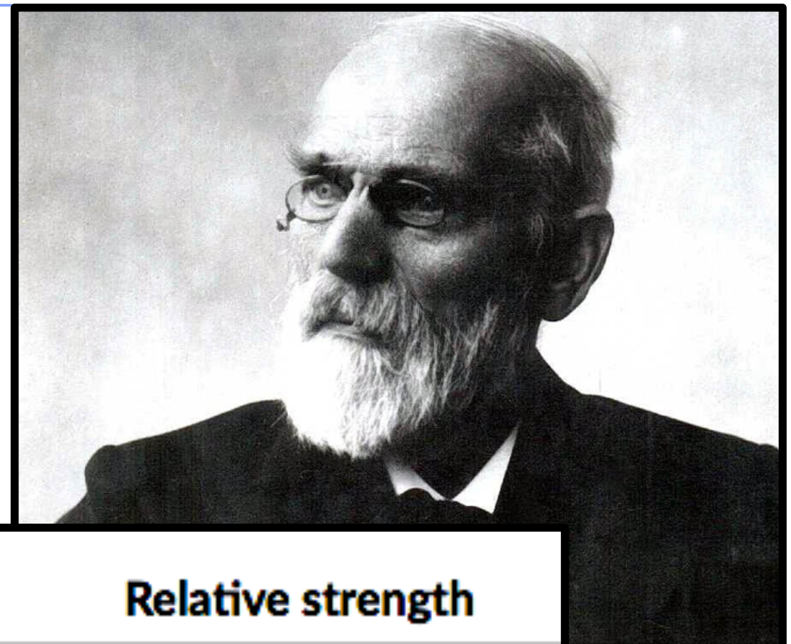


In water, H is bonded directly to O. Thus, water would qualify as having hydrogen bonds between different water molecules in a sample.

Even though acetaldehyde contains both H and O, the bond shown is NOT a hydrogen bond, as the H atom in the above molecule is not directly bonded to the O atom, leaving the H in the top acetaldehyde molecule neutral. This H then is will not attract the partially negatively charged O atom in the lower acetaldehyde molecule!

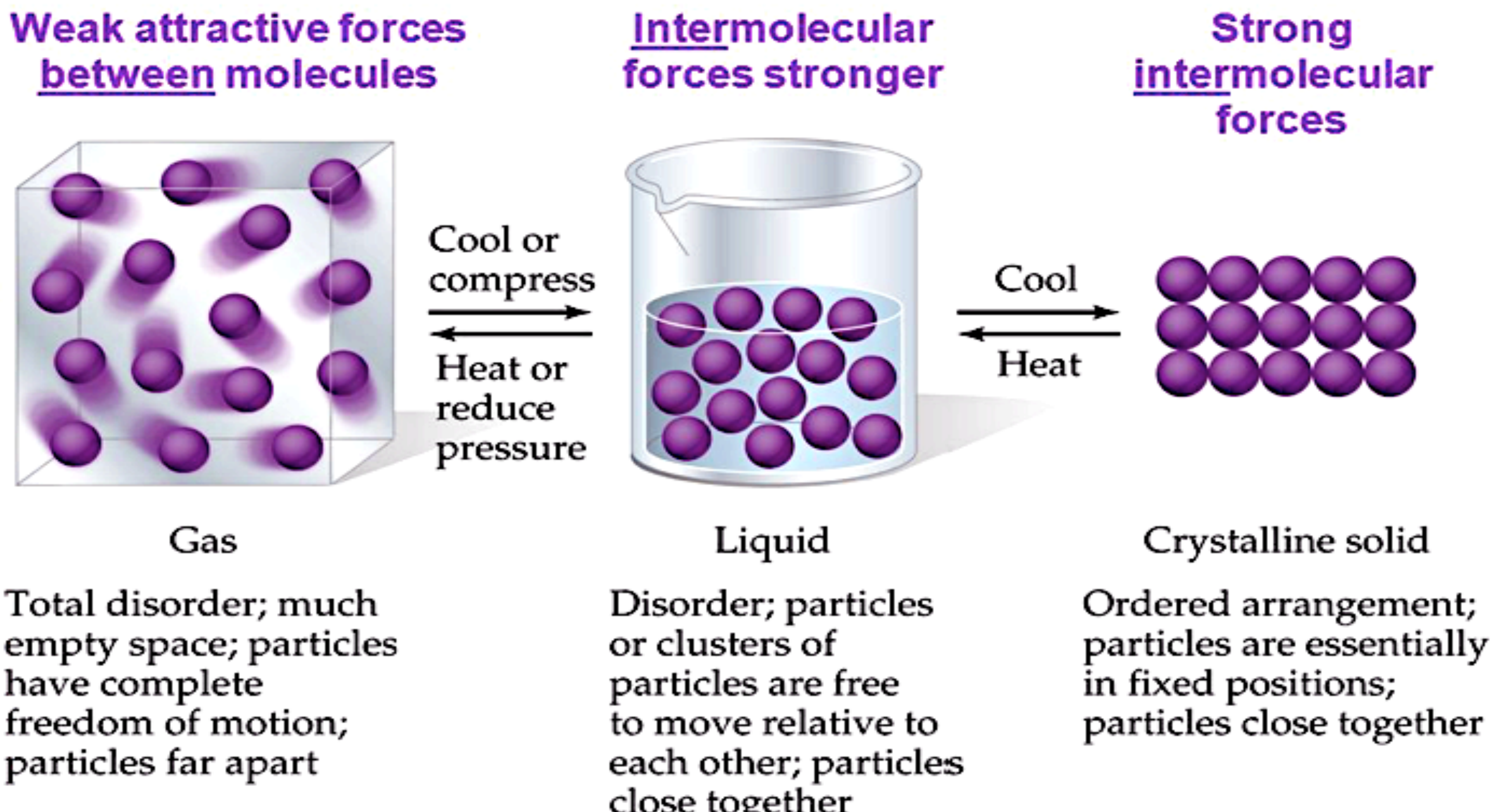
Relative Strength of Van Der Waals Forces





Intermolecular force	Occurs between ...	Relative strength
Dipole-dipole attraction	Partially oppositely charged ions	Strong
Hydrogen bonding	H atom and O, N/ or F atom	Strongest of the dipole-dipole attractions
London dispersion attraction	Temporary or induced dipoles	Weakest

While the existence of any temporary or permanent intermolecular forces cause molecules or ions to attract one another, kinetic energy in the system causes particles to vibrate, collide, and want to disperse in space.



The state of a substance depends largely on the balance between the kinetic energies of the particles and the interparticle energies of attraction.

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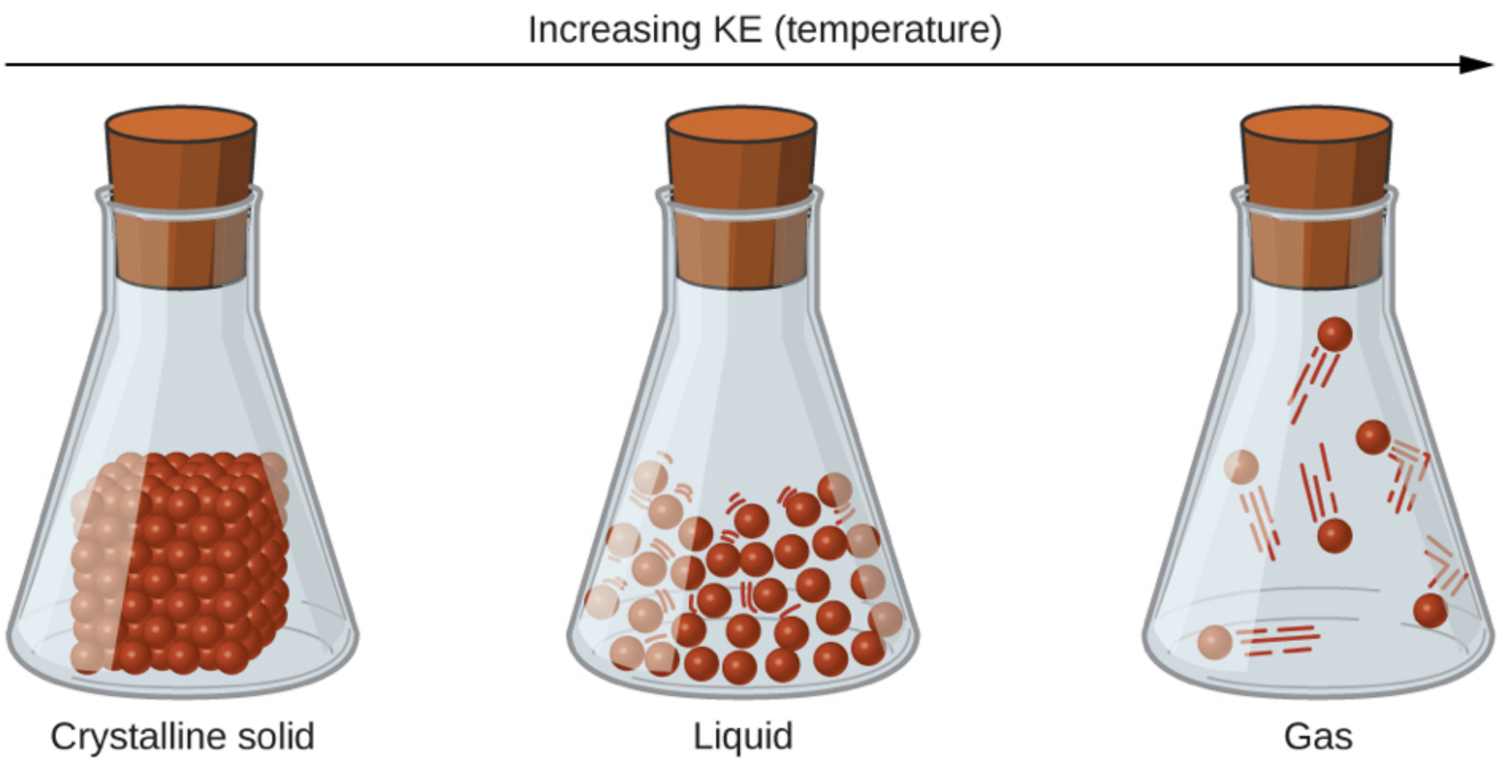
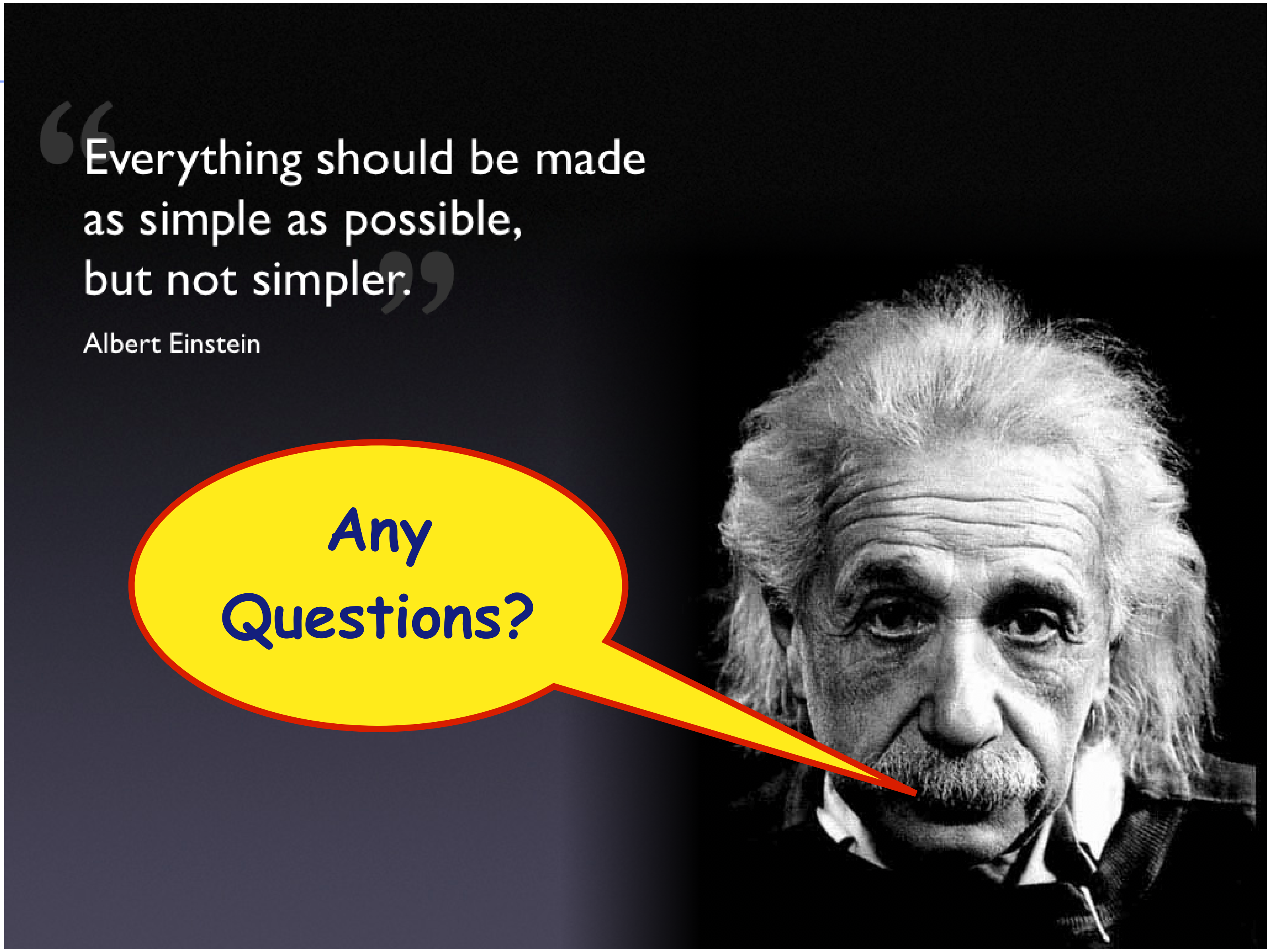
The State of Matter depends on the amount of KE and the strengths of the IMFs present.



Substances that exhibit weaker or occasional, temporary intermolecular forces, change from solid to liquid to gas with the addition of less kinetic energy (at lower temperatures) compared to substances that exhibit stronger or more permanent intermolecular forces as more kinetic energy is needed for the latter to be able to overcome their intermolecular forces of attraction in order to be able to slide past one another in liquid or move far away from one another in gas.


Everything should be made as simple as possible, but not simpler.
Albert Einstein
Any Questions?